9TH CLASS

0333-8033313
راؤ ایاز

0343-7008883
پاکستان زندہ باد

0306-7163117
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STRUCTURE OF ATOMS

Major Concepts:

2.1 Theories and Experiments related to Atomic Structure
2.2 Electronic Configuration
2.3 Isotopes

Time allocation
Teaching periods 16
Assessment periods 03
Weightage 10%

Students Learning Outcomes:

Students will be able to:

- Describe the contributions that Rutherford made to the development of the Atomic Theory.
- Explain how Bohr's atomic theory differed.
- Describe the structure of atom including the location of the proton, electron and neutron.
- Define isotopes.
- Compare isotopes of an atom.
- Discuss the properties of the isotopes of H, C, Cl, U.
- Draw the structure of different isotopes from mass number and atomic number.
- State the importance and uses of isotopes in various fields of life.
- Describe the presence of subshells in shell.
- Distinguish between shells and subshells.
- Write the electronic configuration of first 18 elements in the Periodic Table.

2.1 THEORIES AND EXPERIMENTS RELATED TO STRUCTURE OF ATOM

Q.1 What is the historical background about discovery of atom?

Ans. Ancient Greek Philosopher Democritus suggested that matter is composed of tiny indivisible particles called atoms.

The name atom was derived from Latin Word "Atomos" meaning indivisible.

John Dalton put forward his atomic theory, according to him all the matter is made up of very small indivisible particles called atoms.

In the beginning of 20th century experiments performed by Goldstein, J.J. Thomson, Rutherford, Bohr and other revealed that atom is made up of electron, proton and neutrons and have complicated structure.

Contribution of Dalton: According to Dalton an atom is an indivisible, hard, dense

sphere. Atoms of same element are identical. They combine in different ways to form compounds.

In late 1800's and early 1900's scientists discovered new sub-atomic particles.

Contribution of Goldstein: In 1886, Goldstein discovered positively charged particles called protons.

Contribution of J.J. Thomson: In 1897, J.J. Thomson found in an atom, the negatively charged particles called electrons.

It was established that electrons and protons are fundamental particles of matter.

J.J Thomson (1856-1940) was a British physicist. He was awarded the 1906 Noble Prize in Physics for the discovery of electron and for his work on the conduction of electricity in gases

Plum pudding theory: Thomson put forth his plum pudding theory. He postulated that atoms were solid structures of positively charge with tiny negative particles stuck inside. It is like plums in pudding.

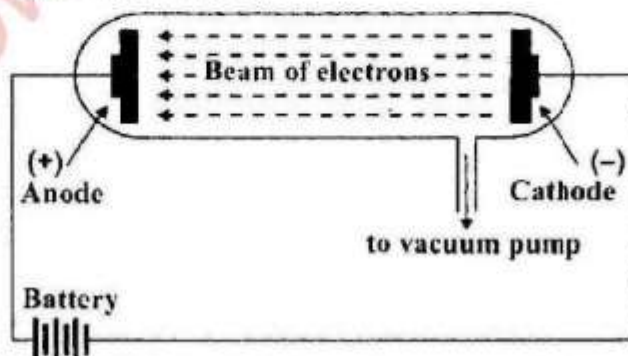
Q.2 Describe the discovery of electron. OR Describe the discharge tube experiment for the discovery of electron.

Ans. Discovery of electron by the passage of electric current through gases at low pressure:

Gases at low pressure:

In 1895, Sir William crooks performed experiments by passing electric current through gases in discharge tube at very low pressure.

He took a glass tube fitted with two metallic electrodes, which were connected to a high voltage battery.

Sir William Crooks (1832-1919) was a British chemist and physicist. He was pioneer of vacuum tubes. He worked on spectroscopy

Discharge tube used for the production of cathode rays.

The pressure inside the tube was kept 10^{-4}atm. When high voltage current was passed through the gas shiny rays were emitted from the cathode and move towards the anode as shown in fig. These rays are called cathode rays because they were originated from the cathode.

Properties of Cathode Rays: Cathode rays have following properties.

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السلام علیکم ورحمۃ اللہ وبرکاتہ:

معزز ممبران: آپ کا وٹس ایپ گروپ ایڈمن "**اردو بکس**" آپ سے مخاطب ہے۔

آپ تمام ممبران سے گزارش ہے کہ:

- گروپ میں صرف PDF کتب پوسٹ کی جاتی ہیں لہذا کتب کے متعلق اپنے کمنٹس / ریویوز ضرور دیں۔ گروپ میں بغیر ایڈمن کی اجازت کے کسی بھی قسم کی (اسلامی وغیر اسلامی، اخلاقی، تحریری) پوسٹ کرنا سختی سے منع ہے۔
- گروپ میں معزز، پڑھے لکھے، سلجھے ہوئے ممبرز موجود ہیں اخلاقیات کی پابندی کریں اور گروپ رولز کو فالو کریں بصورت دیگر معزز ممبرز کی بہتری کی خاطر ریموو کر دیا جائے گا۔
- کوئی بھی ممبر کسی بھی ممبر کو انباکس میں میسج، مس کال، کال نہیں کرے گا۔ رپورٹ پر فوری ریموو کر کے کاروائی عمل میں لائے جائے گی۔
- ہمارے کسی بھی گروپ میں سیاسی و فرقہ واریت کی بحث کی قطعاً کوئی گنجائش نہیں ہے۔
- اگر کسی کو بھی گروپ کے متعلق کسی قسم کی شکایت یا تجویز کی صورت میں ایڈمن سے رابطہ کیجئے۔
- سب سے اہم بات:

گروپ میں کسی بھی قادیانی، مرزائی، احمدی، گستاخِ رسول، گستاخِ امہات المؤمنین، گستاخِ صحابہ و خلفائے راشدین حضرت ابو بکر صدیق، حضرت عمر فاروق، حضرت عثمان غنی، حضرت علی المرتضیٰ، حضرت حسنین کریمین رضوان اللہ تعالیٰ اجمعین، گستاخ اہلبیت یا ایسے غیر مسلم جو اسلام اور پاکستان کے خلاف پراپیگنڈا میں مصروف ہیں یا ان کے روحانی و ذہنی سپورٹرز کے لئے کوئی گنجائش نہیں ہے لہذا ایسے اشخاص بالکل بھی گروپ جوائن کرنے کی زحمت نہ کریں۔ معلوم ہونے پر فوراً ریموو کر دیا جائے گا۔

- تمام کتب انٹرنیٹ سے تلاش / ڈاؤنلوڈ کر کے **فری آف کاسٹ** وٹس ایپ گروپ میں شیئر کی جاتی ہیں۔ جو کتاب نہیں ملتی اس کے لئے معذرت کر لی جاتی ہے۔ جس میں محنت بھی صَرف ہوتی ہے لیکن ہمیں آپ سے صرف دعاؤں کی درخواست ہے۔
- عمران سیریز کے شوقین کیلئے علیحدہ سے عمران سیریز گروپ موجود ہے۔
- **لیڈیز کے لئے الگ گروپ کی سہولت موجود ہے جس کے لئے ویریفکیشن ضروری ہے۔**
- اردو کتب / عمران سیریز یا سٹڈی گروپ میں ایڈ ہونے کے لئے ایڈمن سے وٹس ایپ پر بذریعہ میسج رابطہ کریں اور جواب کا انتظار فرمائیں۔ برائے مہربانی اخلاقیات کا خیال رکھتے ہوئے موبائل پر کال یا ایم ایس کرنے کی کوشش ہر گز نہ کریں۔ ورنہ گروپس سے تو ریموو کیا ہی جائے گا بلاک بھی کیا جائے گا۔

نوٹ: ہمارے کسی گروپ کی کوئی فیس نہیں ہے۔ سب فی سبیل اللہ ہے

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اللہ تبارک تعالیٰ ہم سب کا حامی و ناصر ہو

1. These rays travel in a straight line perpendicular to the cathode surface.
2. They can cast a sharp shadow of an opaque object if placed in their path.
3. Cathode rays are deflected towards positive plate in an electric field showing that they are negatively charged.
4. They raise temperature of the body on which they fall.
5. J.J. Thomson discovered the charge/ mass (e/m) ratio of cathode rays.
6. They produced light when they hit the sides of the discharge tube.
7. It was found that the same type of rays were emitted no matter which gas and which cathode was used in the discharge tube.

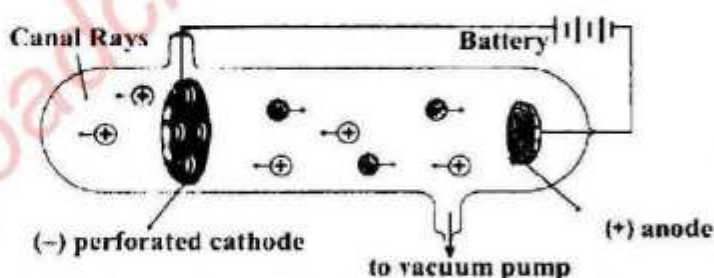
Conclusion: On the basis of these properties it was considered that cathode rays are negatively charged particles called electrons.

Since the nature of cathode rays does not change with the nature of the gas and the cathode used in the tube hence it is safely stated that electrons are the fundamental particles of all atoms.

Q.3 Describe the discovery of proton.

Ans. In 1886 Goldstein discovered that in addition to the cathode rays produced in the discharge tube, other rays were also present.

These rays were traveling in opposite direction to cathode rays. These rays were named as positive rays or anode rays.

He used a discharge tube having perforated cathode as shown in fig.

Discharge tube used for the production of canal rays.

He found that these rays passed through holes present in the cathode and produced a glow on the walls. Hence these are also known as canal rays.

Properties of Anode rays:

1. These rays travel in a straight line in a direction opposite to cathode rays.
2. When electric or magnetic field is applied these rays bend towards the negative pole which shows their positive nature.
3. The nature of positive rays (canal rays) depends upon the nature of gas present in discharge tube.
4. Positive rays do not originate from the anode.

These rays are produced by the collisions of cathode rays (electrons) with the residual gas molecules present in the discharge tube.

$$M + e^- \longrightarrow M^+ + 2e^-$$

The simplest positive rays were obtained when the discharge tube contain

hydrogen gas. These positive rays or hydrogen ions are named as protons. The mass of positively charged particle was found equal to that of a proton or simple multiple of it.

Conclusion: On the basis of these properties it was concluded that like electron, proton is also fundamental particle of an atom.

Q.4 Describe the discovery of neutron.

Ans. Discovery of neutron:

In 1932 James Chadwick discovered neutrons based upon his studies on artificial radioactivity.

When beryllium was bombarded with alpha particles, neutrons were produced.

$$^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ^{1}_{0}n$$

Properties of Neutrons

1. Neutrons are neutral.
2. They are highly penetrating.
3. They are undeflected by electric or magnetic field.
4. Mass of these particles was nearly equal to the mass of proton.

Test yourself 2.1:

i. **Do you know any element having no neutrons in its atoms?**

Ans. **Yes, protium (isotope of hydrogen) has no neutron in it.**

ii. **Who discovered an electron, a proton and a neutron?**

Ans. **Electron: William crooks**
Proton: Gold stein
Neutron: James Chadwick

iii. **How does electron differ from a neutron?**

Ans.

Electron	Neutron
1. Electron is negatively charged	1. Neutron is neutral
2. Electrons are revolving around the nucleus	2. Neutrons are present in the nucles of an atom

iv. **Explain, how anode rays are formed from the gas present in the discharge tube?**

Ans. **Anode rays are produced in the gas discharge tube due to the collision of cathode rays with the residual gas present in the discharge tube.**

Q.5 (a) Describe the Rutherford's experiment for the discovery of nucleus. Explain his atomic model. What are defects in this model.
(b) Write down the objections of scientists on Rutherford atomic model.

Ans: Rutherford experiment for the discovery of Nucleus:

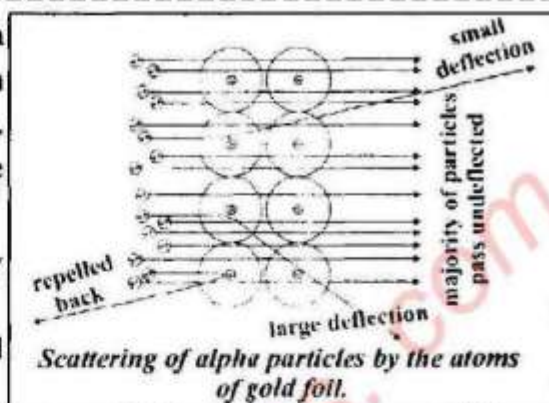
In order to determine the structure of atom, Rutherford carried out an experiment in year 1911.

He bombarded a very thin (0.00004cm thick) gold foil with alpha particles from a radioactive source (Radium or polonium). Alpha particles are Helium nuclei (He^{+2}) which can penetrate through matter to some extent.

Rutherford observed the effects of alpha particles on a photographic plate or a screen coated with zinc sulphide as shown in figure. He proved that the plum-pudding model of the atom was not correct.

Scattering of alpha particles by the atoms of gold foil.

Observations: Observations made by Rutherford were as follows.

1. Most of the alpha particles passed through the gold foil un-deflected.
2. Out of 20000 particles, only a few were deflected at fairly large angles and very few bounced back on hitting the gold foil.

Results of the experiment:

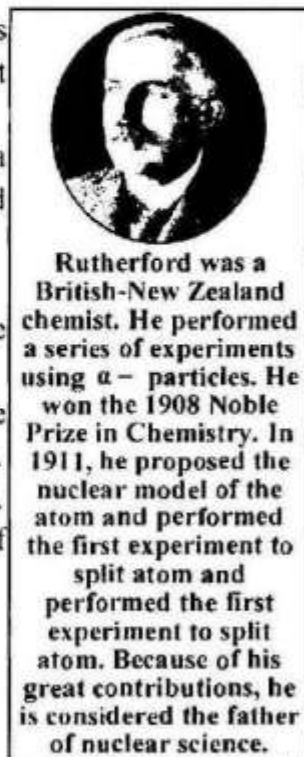
Keeping in view the experiment, Rutherford proposed planetary model for an atom and concluded following results.

1. Rutherford concluded that as most of the alpha particles went through the gold foil undeflected, it means major part of an atom is empty.
2. The deflection of a few particles proved that there is a center of positive charges in an atom which is called nucleus of an atom.
3. Nucleus is located at the center of the atom.
4. The complete rebounce of a few particles shows that the nucleus is very dense and hard.
5. Since a few particles were deflected it shows that size of the nucleus is very small as compared to the volume of an atom.
6. The whole mass of the atoms in concentrated in the nucleus.
7. An atom as a whole is neutral, therefore the number of electrons in an atom in equal to the number of protons.
8. Except electrons, all other fundamental particles that lie within a nucleus are called nucleons.
9. The electrons revolve around the nucleus.

Rutherford was a British-New Zealand chemist. He performed a series of experiments using α – particles. He won the 1908 Noble Prize in Chemistry. In 1911, he proposed the nuclear model of the atom and performed the first experiment to split atom and performed the first experiment to split atom. Because of his great contributions, he is considered the father of nuclear science.

Defects in Rutherford's Model:

Rutherford's experiment proved that the plum-pudding model of an atom was not correct yet it had following defects.

1. According to classical theory of radiation, electrons being the charged particles should release or emit energy continuously and they should ultimately fall into the nucleus.
2. If the electrons emit energy continuously, they should form a continuous spectrum

but infact line spectrum was observed.

(b) Objections of scientists on Rutherford's atomic model.

The scientists had objections on Rutherford atomic model. They initiated the quest to answer they following questions.

1. How can an atom collapse or why are atoms stable?
2. Why does an atom give line spectrum?
3. Scientists considered there must be another model of atom. It indicated that Rutherford's model was not perfect.

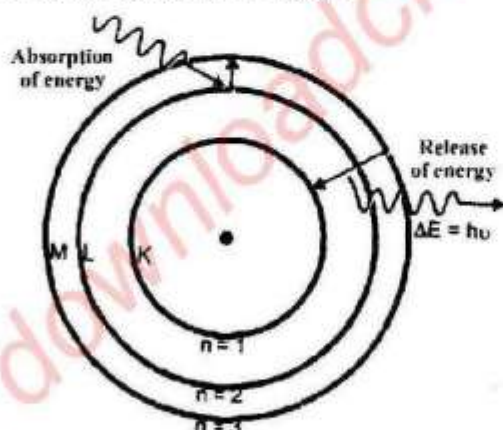
Q.6 Explain the Bohr's atomic model.

Ans. Bohr's atomic model:

Neil Bohr presented a new model for the structure of atom in 1913.

The Quantum Theory of Max planck was used as foundation for this model.

According to Bohr's atomic model revolving electron in an atom does not absorb or emit energy continuously.

The energy of revolving electron is quantized, as it revolves only in orbits of fixed energy celled energy levels by him. The Bohr's atomic model is shown in figure.

Bohr's atomic model showing orbits.

Neil Bohr was a Danish physicist who joined Rutherford in 1912 for his post doctoral research. In 1913, Bohr presented his atomic model based upon Quantum theory. He won the 1922 Noble Prize for Physics for his work on the structure of an atoms.

Main Postulates of Bohr's atomic Model:

The main postulates of Bohr's atomic model are following.

1. The hydrogen atom consists of a tiny nucleus and electrons are revolving in one of circular orbits of radius "r" around the nucleus.
2. Each orbit has a fixed energy that is quantized.
3. As long as the electron revolves in a fixed orbit, its energy remains constant. The energy is emitted or absorbed only when an electron jumps from one orbit to another.
4. When an electron jumps from lower orbit to higher orbit, it absorbs energy and when it jumps back from higher orbit to lower orbit it radiates energy. This change

in energy Δ E is given by following Planck's equation.

$$\Delta E = E_2 - E_1 = h\upsilon$$

$$\Delta E = h\upsilon$$

Where 'h' is Planck constant and its value is 6.63×10^{-34} Js and υ is the frequency of light.

5. According to Bohr's atomic model, electron can revolve only in orbits of fixed angular momentum *mvr*, given as:

$$mvr = n\frac{h}{2\pi}$$

Where "n" is the quantum number or orbit number having values 1,2,3 and so on.

Do you know?

Quantum means fixed energy. It is the smallest amount of energy that can be emitted or absorbed as electromagnetic radiation. Quanta is plural of quantum.

In 1918 Noble prize in physics was awarded to German physicist Max Planck (1858-1947) for his work on the quantum theory.

Q.7 Write down the differences between Rutherford atomic theory and Bohr's atomic theory.

Ans.

	Rutherford's Atomic Theory	Bohr's Atomic Theory
(i)	It was based upon classical theory.	It was based upon quantum theory.
(ii)	Electrons revolve around the nucleus.	Electrons revolve around the nucleus in orbits of fixed energy.
(iii)	No idea about orbits was introduced.	Orbits had angular momentum.
(iv)	Atoms should produce continuous spectrum.	Atoms should produce line spectrum.
(v)	Atoms should collapse.	Atoms should exist.

Test yourself 2.2:

(i) How was it proved that the whole mass of an atom is located at its centre?

Ans. Rutherford proved that most of the alpha particles pass through the gold foil which proves that most of the volume occupied by an atom is empty. Only few particles are bounced back, which shows that whole the mass of atom is located at the center of atom.

(ii) How was it shown that atomic nuclei are positively charged?

Ans. Rutherford experiment proved that a few alpha particles bounced back at various angles which proves that there is a center of positive charges in an atom.

(iii) Name the particles which determine the mass of an atom.

Ans. Protons and neutrons

(iv) What is the classical theory of radiation? How does it differ from quantum theory?

Ans. According to classical theory electrons, being charged particles should release or emit energy continuously and they should ultimately fall into the nucleus.

According to quantum theory as long as on electron remains in a particular orbit it does not emit or absorb energy. The energy is emitted or absorbed only when an electron jumps from one orbit to another.

(v) How can you prove that angular momentum is quantized?

Hint: Let angular momentum (mvr) of 1st orbit is mvr = nh/2 π
By putting the values of h and π

$$mvr = \frac{6.63\times10^{-34}}{2\times3.14} = 1.0\times10^{-34}\ kg\ m^2 s^{-1}$$

Ans. $mvr = \frac{nh}{2\pi}$

By putting the values of "h" and π

$$mvr = \frac{1\times6.63\times10^{-34}}{2\times3.14} = 1.0\times10^{-34}\ kg\ m^2 s^{-1}$$

Angular momentum of second orbit

$$= \frac{2\times6.63\times10^{-34}}{2\times3.14} = 2.1\times10^{-34}\ kg\ m^2 s^{-1}$$

Angular momentum of third orbit

$$= \frac{3\times6.63\times10^{-34}}{2\times3.14} = 3.18\times10^{-34}\ kg\ m^2 s^{-1}$$

Which proves that only specific values of angular momentum are possible.

2.2 ELECTRONIC CONFIGURATION

Q.8 What is meant by shell (energy level) and subshell.

Ans. Shell or energy level:

There is a probability of finding the electrons in certain regions of space around the nucleus of the atom. This region of the space is called shell or energy level.

Sub-shell: Each shell has further sub energy level in it. These sub energy levels are called sub-shell.

Types of sub-shell: There are four types of sub-shells.

(i) s - sub-shell (ii) p - sub-shell (iii) d - sub-shell (iv) f - sub-shell

Number of electrons in Sub-shells

Sub-shell	Maximum electrons
s	2
p	6
d	10
f	14

Number of Sub-shells in Various Shells

n value	Shell	Sub-shell
1	K	only s
2	L	s, p
3	M	s, p, d
4	N	s, p, d, f

Q.9 What is electronic configuration? How electrons are arranged in shells?

Ans. The arrangement of electrons around the nucleus is called electronic configuration. The maximum number of electrons which can accommodate in an energy level (shell) is given by formula.

"$2n^2$"

Where "n" represents the energy level or shell.

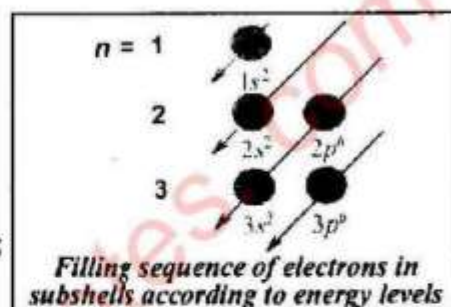
According to formula.

Number of electrons in K-shell = $2n^2$, $2(1)^2 = 2$

Number of electron in L-shell = $2n^2 = 2(2)^2 = 8$

Number of electron in M-shell = $\propto n^2 = 2(3)^2 = 18$

Number of electron in N-shell = $2n^2 = 2(4)^2 = 32$

Filling sequence of electrons in subshells according to energy levels

Q.10 Write down the electronic configurations of following.

(i) Sodium (ii) Argon (iii) Sulphur (iv) Chlorine

Ans. (i) Electronic Configuration Sodium $\left(Na^{23}_{11}\right)$

Atomic number of sodium = 11

Number of electrons in sodium = 11

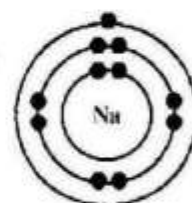
Distribution of electrons in sub-shells

$1s^2$, $2s^2\ 2p^6$, $3s^1$

Distribution of electrons in shells

Number of electrons in K-shell = 2

Number of electron in L-shell = 8

Number of electron in M-shell = 1

(ii) Electronic Configuration of Argon $\left(\begin{matrix}36\\ Ar\\ 18\end{matrix}\right)$

Atomic Number of Argon = 18

Number of electron in Argon = 18

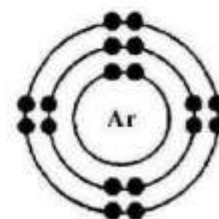
Distribution of electrons in sub-shells

$1s^2$, $2s^2\ 2p^6$, $3s^2$, $3p^6$

Distribution of electrons in shells

Number of elecrons in K-shell = 2

Number of electronic L-shell = 8

Number of electron in M-shell = 8

(iii) Electronic configuration of Sulphur, $\left(\begin{matrix}32\\ S\\ 16\end{matrix}\right)$

Atomic Number of sulphur = 16

Number of eletrons in sulphur = 16

Distribution of electrons in sub-shells

$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^4$

Distribution of electrons in shells

Number of electrons in K-shell = 2
Number of electrons in L-shell = 8
Number of electrons in M-shell= 6

(iv) Electronic configuration of Chloride ion (Cl^-)

Atomic Number of chlorine = 17
Number of electrons in chloride ion = 17+1=18

Distribution of electrons in sub-shells

$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^5$

Distribution of electrons in shells

Number of electrons in K-shell = 2
Number of electrons in L-shell = 8
Number of electrons in M-shell= 8

Example 2.1

Write the electronic configuration of an element having 11 electrons.

Solution:

Keep in mind that all electrons do not have the same energy. Therefore, they are accommodated in different shells according to increasing energy and capacity of the shell. First of all, the electrons will go to K shell which has minimum energy. It can accommodate 2 electrons. After this, electrons will go to L shell that can accommodate 8 electrons. Thus K and L shells have accommodated 10 electrons. The remaining 1 electron will go to M shell, the outermost shell of maximum energy in this case.

The electronic configuration will written as: K L M
2, 8, 1,

But it is not necessary to write the subshells. Therefore, it is simply written as 2,8, and 1. Further distribution of electrons in subshells will be: $1s^2$, $2s^2$, $2p^6$, $3s^1$.

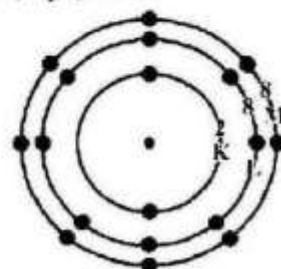
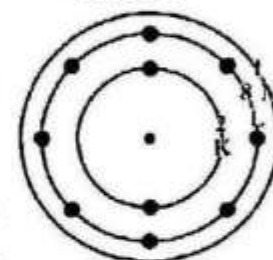
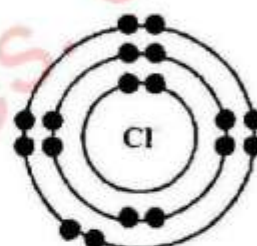
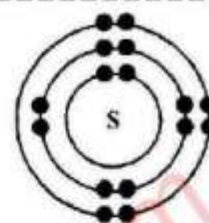
Example 2.2

Write down the electronic configuration of Cl^- ion

Solution:

We know that chlorine has 17 electrons and chloride ion (Cl^-) has 17+1=18 electrons. Its electronic configuration will be 2, 8, 8, which is presented in the figure. The further distribution of electrons in subshells will be $1s^2$, $2s^2$, $2p^6$, $3s^2$ $3p^6$.

Example 2.3

An element has 5 electrons in M shell. Find out its atomic number.

Solution:

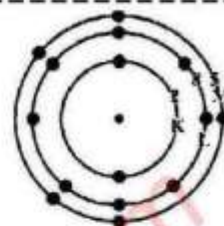
When there are 5 electrons in M shell, it means K and L shell are completely filled with their maximum capacity of 10 electrons. Hence the electronic configuration of the element is:

K L M

2, 8, 5, or just 2, 8, 5

So, the total number of electrons is 2 + 8 + 5 = 15

As we know, the number of electrons in an atom is equal to its atomic number.

Therefore, atomic number of this element is 15.

Q.11 Make a table which shows electronic configuration of first 18 elements.

Ans. Electronic Configuration of First Eighteen Elements

Element	Symbol	Atomic Number	Electronic Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2, 2s^1$
Beryllium	Be	4	$1s^2, 2s^2$
Boron	B	5	$1s^2, 2s^2, 2p^1$
Carbon	C	6	$1s^2, 2s^2, 2p^2$
Nitrogen	N	7	$1s^2, 2s^2, 2p^3$
Oxygen	O	8	$1s^2, 2s^2, 2p^4$
Fluorine	F	9	$1s^2, 2s^2, 2p^5$
Neon	Ne	10	$1s^2, 2s^2, 2p^6$
Sodium	Na	11	$1s^2, 2s^2, 2p^6, 3s^1$
Magnesium	Mg	12	$1s^2, 2s^2, 2p^6, 3s^2$
Aluminium	Al	13	$1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
Silicon	Si	14	$1s^2, 2s^2, 2p^6, 3s^2, 3p^2$
Phosphorus	P	15	$1s^2, 2s^2, 2p^6, 3s^2, 3p^3$
Sulphur	S	16	$1s^2, 2s^2, 2p^6, 3s^2, 3p^4$
Chlorine	Cl	17	$1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
Argon	Ar	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Test yourself 2.3:

(i) How many the maximum number of electrons that can be accommodated in a p-subshell?

Ans. Six electrons can be accomodated in a p-subshell.

(ii) How many subshells are there in second shell?

Ans. Two "s" and p subshells are there in second shell.

(iii) Why does an electron first fill 2p orbital and then 3s orbital?

Ans. Electron first fill 2p orbital then 3s because the energy of 2p orbital is less than 3s.

(iv) If both K and L shells of an atom are completely filled; How many total number of electrons are present in them?

Ans. Ten (10)

(v) How many electrons can be accommodated in M shell?

Ans. 18 electrons can be accommodated in M shell.

(vi) What is the electronic configuration of a hydrogen atom?

Ans. Electronic configuration of hydrogen atom = $1s^1$

(vii) What is atomic number of phosphorus? Write down its electronic configuration.

Ans. Atomic number of phosphorous = 15

Electronic configuration of phosphorous = $1s^2 2s^2 2p^6 3s^2 3p^3$

(viii) If an element has atomic number 13 and atomic mass 27; how many electrons are there in each atom of the element?

Ans. Number of electrons = 13

(ix) How many electrons will be in M shell of an atom having atomic number 15,

Ans. Number of electrons in "M" of an atom having atomic Number 15 is = 5

(x) What is maximum capacity of a shell?

Ans. The maximum capacity of a shell can be determined by using formula "$2x^2$" where "n" shows the number of orbit.

2.3 ISOTOPES

Q12. What is meant by isotopes? Explain with examples.

Ans. Isotopes:

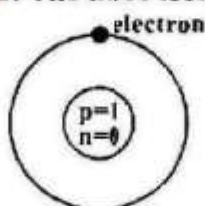
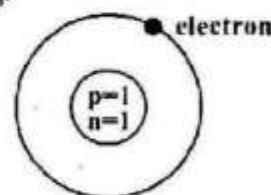
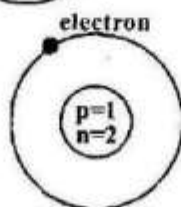
Atoms of same element having same atomic number but different atomic masses are called isotopes. **OR**

Atoms of same element having different number of neutrons are called isotopes.

Examples

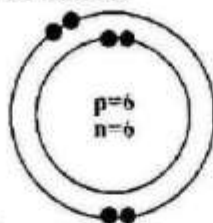
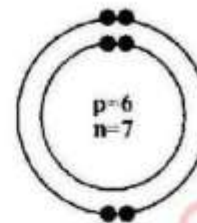
Isotopes of Hydrogen:

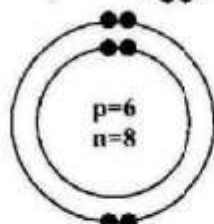
The naturally occurring hydrogen is combination of its three isotopes, present in different abundances. The three isotopes of hydrogen are following.

(i) Protium H_1^1

(ii) Deuterium H_1^2 or D_1^2

(iii) Tritium H_1^3

(ii) Isotopes of Carbon:

There are two stable isotopes of carbon ^{12}C and ^{13}C and one radioactive isotope ^{14}C. The isotope ^{12}C is present in abundance of 98.9% while ^{13}C and ^{14}C are both

present in only 1.1% in nature.

(i)

(ii) C_6^{13}

(iii) C_6^{14}

(iii) Isotopes of Uranium:

There are three isotopes of Uranium.

(i) $^{234}_{92}U$ (ii) $^{235}_{92}U$ (iii) $^{238}_{92}U$

The $^{238}_{92}U$ is found in nature nearly 99% pure.

All the isotopes of an element occupy same position in the periodic table.

Isotopes of Chlorine:

There are two isotopes of chlorine.

(i) $^{35}_{17}Cl$ (ii) $^{37}_{17}Cl$

Atomic Number, Mass Number, Number of Protons and Neutrons of H, C, Cl and U

Symbol	Atomic Number	Mass Number	No. of Proton	No. of Neutron
^{1}H	1	1	1	0
^{2}H	1	2	1	1
^{3}H	1	3	1	2
^{12}C	6	12	6	6
^{13}C	6	13	6	7
^{14}C	6	14	6	8
^{35}Cl	17	35	17	18
^{37}Cl	17	37	17	20
^{234}U	92	234	92	142
^{235}U	92	235	92	143
^{238}U	92	238	92	146

Science, Technology, Society:

Application of Isotopes: **In science and many different technological fields, isotopes have vast applications. The biggest application is in the field of medicine. They are applied in diagnosis, radiotherapy and treatment of many diseases like cancer.**

Q13. Write down the applications (uses) of isotopes.

Ans. Applications of isotopes:

(i) Radiotherapy (Treatment of cancer):

For the treatment of cancer (skin cancer) isotopes like P-32 and Sr – 90 are used because they emit less penetrating Beta radiations.

Co – 60 affecting within the body, is used because it emits strongly penetrating gamma rays.

(ii) Tracer for diagnosis and Medicine

The radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body.

Isotopes of iodine - 131 are used for diagnosis of goiter, in thyroid gland. Technetium is used to monitor the bone growth.

(iii) Archaeological and Geological uses:

The radioactive isotopes are used to estimate the age of fossils like dead plants and animals and stones etc.

The age determination of very old objects based on the half-lives of the radioactive isotope called radioactive isotope dating.

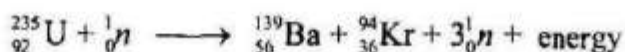
An important method of age determination of old carbon containing objects (fossils) by measuring the radioactivity of C-14, in them is called radioactive dating or simply carbon dating.

(iv) Chemical reaction and structure determination:

The radioactive isotopes are used in a chemical reaction to follow a radioactive element during the reaction and ultimately to determine the structure e.g. C-14 is used to label CO_2. As CO_2 is used by plants for Photosynthesis to form glucose, its movement is detected through the various intermediate steps up to glucose.

(v) Applications in Power generation:

The radioactive isotopes are used to generate electricity by carrying out controlled nuclear fission reactions in nuclear reactors. e.g. U-235 is bombarded with slow moving neutrons, the uranium breaks to produce Barium-139 and Krypton - 94 and three neutrons.

$$^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n + \text{energy}$$

A large amount of energy is released which is used to convert water into steam boilers. The steam drives the turbine to generate electricity. This is the peaceful use of atomic energy.

Test yourself 2.4:

(i) Why do the isotopes of an element have different atomic masses?

Ans. The isotopes of an element have different atomic masses due to the presence of different number of neutrons.

(ii) How many neutrons are present in C-12 and C-13?

Ans. Number of neutrons in C-12 = 6
Number of neutrons in C-13 = 7

(iii) Which of the isotopes of hydrogen contains greater number of neutrons?

Ans. Tritium contains greater number of neutrons (2).

(iv) Give one example each of the use of radioactive isotope in medicine and radiotherapy.

Ans. **Use of radioactive isotope in medicines:** The radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body. Isotopes of iodine -131 are used for diagnosis of goiter in thyroid gland.

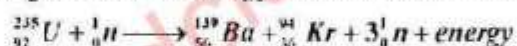
Use of radioactive isotope in radiotherapy: Isotopes P-32 and Sr-90 are used for the treatment of skin cancer.

(v) How is the goiter in thyroid gland detected?

Ans. Isotopes of iodine -131 are used for diagnosis of goiter in thyroid.

(vi) Define nuclear fission reaction.

Ans. **Nuclear fission reaction:** When U-235 is bombarded with slow moving neutrons, the uranium nucleus breaks up to produce Barium-139 and Krypton-94 and three neutrons with evolution of a large amount of energy. This is called nuclear fission reaction.

$$^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n + energy$$

(vii) When U-235 breaks up, it produces a large amount of energy. How is this energy used?

Ans. The energy produced is used to convert water into steam in boiler. The steam drives the tribunes to generate electricity.

(viii) How many neutrons are produced in the fission reaction of U-235?

Ans. Three neutrons are produced in the fission reaction of U-235.

(ix) U-235 fission produces two atoms of which elements?

Ans. *(i) Barium - 139* *(ii) Krypton - 94*

Science, Technology, Society:

Testing Prevailing Theories Brings About Change in Them: Science is a process for producing knowledge. The process depends both on making careful observations of phenomenae and inventing theories for making sense out of those observations. Change in knowledge is inevitable because new observations may challenge prevailing theories. No matter how well one theory explains a set of observations, it is possible that another theory may fit just as well or better, or may fit a still wider range of observations. In science, the testing and improving and occasional discarding of theories, whether new or old, go on all the time. Scientists assume that even if there is no way to secure complete and absolute truth, increasingly accurate approximations can be made to account for the world and how it works.

Key Points

- Cathode rays were discovered in last decade of nineteen century. The properties of cathode rays were determined and they led to the discovery of electron.
- Canal rays were discovered in 1886 by Goldstein. The properties of canal rays

resulted in the discovery of proton in the atom.

- Neutron in the atom was discovered in 1932 by Chadwick.
- First of all structure of an atom was presented by Rutherford in 1911. he proposed that an atom contains nucleus at the centre and electrons revolve around this nucleus.
- Bohr presented an improved model of an atom in 1913 based upon four postulates. He introduced the concept of circular orbit, in which electrons revolve. As long as electron remains in a particular orbit, it does not radiate energy. release and gain of energy is because of change of orbit.
- The concept of shells and subshells is explained.
- A shell consists of subshells.
- Isotopes are defined as the atoms of elements that have the same atomic number but different atomic mass.
- Hydrogen, carbon and uranium have three isotopes each, whereas chlorine has two isotopes.

Exercise (Solved)

Multiple Choice Questions

Put a (✓) on the correct answer.

1. Which one of the following results in the discovery of proton?
(a) cathode rays (b) canal rays (c) X-rays (d) alpha rays

2. Which one of the following is the most penetrating?
(a) protons (b) electrons (c) neutrons (d) alpha particles

3. The concept of orbit was used by
(a) J.J. Thomson (b) Rutherford (c) Bohr (d) Planck

4. Which one of the following shell consists of three subshells?
(a) O shell (b) N shell (c) L shell (d) M shell

5. Which radioisotope is used for the diagnosis of tumor in the body?
(a) cobalt-60 (b) iodine-131 (c) strontium-90 (d) phosphorus-32

6. When U-235 breaks up, it produces:
(a) electrons (b) neutrons (c) protons (d) nothing

7. The p subshell has:
(a) one orbital (b) two orbitals (c) three orbitals (d) four orbitals

8. Deuterium is used to make:
(a) light water (b) heavy water (c) soft water (d) hard water

9. The isotope C-12 is present in abundance of:
(a) 96.9% (b) 97.6% (c) 99.7% (d) none of these

10. Who discovered the proton?
(a) Goldstein (b) J.J Thomson (c) Neil Bohr (d) Rutherford

Answers: **1.** canal rays **2.** neutrons **3.** Bohr **4.** M shell
5. iodine-131 **6.** neutrons **7.** three orbitals **8.** heavy water
9. none of these **10.** Goldstein

Short Answer Questions.

1. What is the nature of charge on cathode rays?

Ans. Negative

2. Give five characteristics of cathode rays.

Ans. For answer see Q. 2.

3. The atomic symbol of a phosphorus ion is given as $^{31}_{15}P^{3-}$.

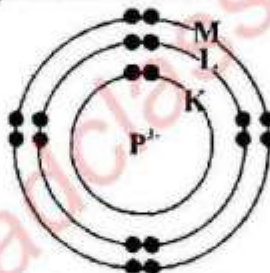
(a) How many protons, electrons and neutrons are there in the ion?

Ans. Proton = 15, Electron = 18, Neutrons = 31-15 = 16

(b) What is name of the ion?

Ans. Phosphide ion.

(c) Draw the electronic configuration of the ion.

Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

(d) Name the noble gas which has the same electronic configuration as the phosphorus ion has.

Ans. Argon.

4. Differentiate between shell and subshell with examples of each.

Ans. For Answer See Q. 8

5. An element has an atomic number 17. How many electrons are present in K, L and M shells of the atom?

Ans. Number of electrons in K-shell = 2 , Number of electrons in L-shell = 8
Number of electrons in M-shell = 7

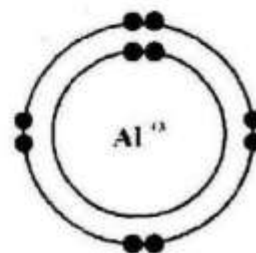
6. Write down the electronic configuration of Al^{3+}. How many electrons are present in its outermost shell?

Ans. Electronic configuration of Al^{3+}
Number of electrons in Al^{+3} = 10

Distribution of electrons in sub-shells

$1s^2, 2s^2, 2p^6$

Distribution of electrons in shells

Number of electrons in K-shell = 2
Number of electrons in L-shell = 8
Number of elecrons in the outermost shell of Al^{+3} = 8

7. Magnesium has electronic configuration 2, 8, 2,

(a) How many electrons are in the outermost shell?

Ans. 2

(b) In which subshell of the outermost shell electrons are present?

Ans. Electronic configuration = $1s^2, 2s^2, p^6, 3s^2$

The outermost electrons are present in 3s. sub-shells.

(c) Why magnesium tends to lose electrons?

Ans. Magnesium loses two electrons in order to get stable electronic configuration.

8. What will be the nature of charge on an atom when it loses an electron or when it gains an electron?

Ans. When an atom loses an electron it becomes positively charged. When an atom gains an electron it becomes negatively charged.

9. For what purpose U-235 is used?

Ans. U-235 is used to generate electricity which is the peaceful use of atomic energy.

10. A patient has goiter. How will it be detected?

Ans. The goiter can be detected by using iodine - 131 isotopes.

11. Give three properties of positive rays.

Ans. For answer See Q. 3

12. What are the defects of Rutherford's atomic model?

Ans. For answer See Q. 5(a)

13. As long as electron remains in an orbit, it does not emit or absorb energy. When does it emit or absorb energy?

Ans. The energy is emitted or absorbed only when an electrons jumps from one orbit to another.

Long Answer Questions

Q.1 How are cathode rays produced? What are their five major characteristics?

Ans. For answer see Q. 2.

Q.2 How was it proved that electrons are fundamental particles of an atom?

Ans. For answer see Q. 2.

Q.3 Draw a labeled diagram to show the presence of protons in the discharge tube and explain how were canal rays produced?

Ans. For answer see Q. 3.

Q.4 How did Rutherford discovered that atom has a nucleus located at the centre of the atom?

Ans. For answer see Q. 5(a).

Q.5 One of the postulates of Bohr's atomic model is that angular momentum of a moving electron is quantized. Explain its meaning

and calculate the angular momentum of third orbit (i.e. n=3)

Ans. The angular momentum of electron is fixed and only these orbits are possible in which angular momentum of electron is $mvr = n\frac{h}{2\pi}$

If $n = 3$

$$mvr = 3\frac{h}{2\pi}$$

Q.6 How did Bohr prove that an atom must exist?

Ans. For answer see Q. 6.

Q.7 What do you mean by electronic configuration? What are basic requirements while writing electronic configuration of an element (atom)?

Ans. For answer see Q. 9.

Q.8 Describe the electronic configuration of Na^+, Mg^{2+} and Al^{3+} ions. Do they have the same number of electrons in the outermost shell?

Ans. Electronic Configuration of Na^+:

Distribution of Electrons in sub-shells: Number of electrons in $Na^+ = 10$

Distribution of electrons in sub-shells: $1s^2, 2s^2, 2p^6$

Distribution of electrons in shells:

Number of electrons in K-shell = 2

Number of electrons L-shell = 8

Electronic Configuration of Mg^{+2}:

Number of electrons in $Mg^{+2} = 10$

Distribution of Electrons in sub-shells:

$1s^2, 2s^2, 2p^6$

Number of electrons in shells:

Number of electrons in K-shell = 2

Number of electrons in L-shell = 8

Electronic configuration of Al^{3+}:

Number of electrons in $Al^{+3} = 10$

Distribution of electrons in sub-shells:

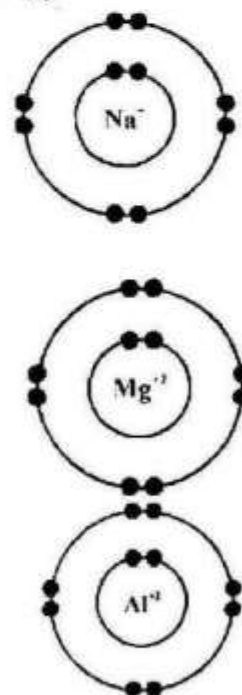
$1s^2, 2s^2, 2p^6$

Yes: They have same number of elecrons in the outermost shell.

Q.9 Give the applications of isotopes in the field of radiotherapy and medicines.

Ans. For answer see Q. 12.

Q.10 What is an isotope? Describe the isotopes of hydrogen with diagrams.

Ans. For answer see Q. 11. (Only isotopes of Hydrogen).

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

2.1 Theories And Experiments Related To Structure Of Atom

2.2 Electronic Configuration

☆ **Tick the correct answer.**

1. **The mass of electron is:** (LHR. GI, MLN. GII)
   (A) 9.106×10^{-28}g (B) 1.674×10^{-24}g (C) 1.672×10^{-24}g (D) 1.66×10^{-24}g
2. **The nucleus of an atom is composed of:** (LHR. GI, MLN. GII)
   (A) Electrons (B) Electrons and protons
   (C) Electrons and neutrons (D) Protons and neutrons
3. **Who discovered proton?** (GRW. GI, RWP. GII, BWP. GII, FBD. GII, SWL. GI)
   (A) Goldstein (B) Rutherford (C) Chadwick (D) Bohr
4. **Mass of Neutron is:** (SGD. GII)
   (A) 1.0073 amu (B) 1.0080 amu (C) 1.0087 amu (D) 1.0097 amu
5. **The concept of orbit of atom was introduced by:** (RWP. GI, FBD. GI)
   (A) J.J. Thomson (B) Rutherford (C) Bohr (D) Plancks
6. **Who discovered Cathode Rays:** (LHR. GI, BWP. GII)
   (A) Gold Stein (B) John Dalton
   (C) Sir William Crooks (D) Neil Bohr
7. **Neutron was discovered by:** (MLN. GI)
   (A) Crooks (B) Bohr (C) Rutherford (D) Chadwick
8. **In discharge tube the canal rays are produced due to:** (MLN. GII)
   (A) Presence of Anode (B) Due to the ionization of gas molecules
   (C) Presence of Cathode (D) Due to high pressure of gas
9. **Who is the Father of Nuclear Science?** (SGD. GI)
   (A) Neil Bohr (B) Rutherford (C) Max Planck (D) J.J Thomson
10. **Cathode rays have charge:** (RWP. GII)
    (A) negative (B) positive (C) neutral (D) ionic bond
11. **Which one of the following particle is most penetrating:** (BWP. GI)
    (A) proton (B) electron (C) neutron (D) alpha particle
12. **How much electrons can be accommodated in M-Shell:** (LHR. GII)
    (A) 8 (B) 18 (C) 20 (D) 40

13. How many electrons can K-shell accommodate? (GRW. GII, BWP. GII)

(A) 3 (B) 2 (C) 4 (D) 5

14. The atomic number of fluorine is: (FBD. GII)

(A) 3 (B) 4 (C) 5 (D) 9

15. __________ shell consists of three subshells. (MLN. GI, DGK. GI)

(A) O Shell (B) N Shell (C) L Shell (D) M Shell

16. "N" Shell can accommodate electrons: (SGD. GI)

(A) 18 (B) 32 (C) 8 (D) 2

17. Electronic configuration of Nitrogen is: (SGD. GII)

(A) $1s^2, 2s^2, 2p^2$ (B) $1s^2, 2s^2, 2p^3$

(C) $1s^2, 2s^2, 2p^4$ (D) $1s^2, 2s^2, 2p^5$

18. After gaining one electron, chlorine atom attains the electronic configuration of which noble gas: (LHR. GII)

(A) Helium (B) Neon (C) Argon (D) Krypton

19. Which shell consists of four sub-shells? (GRW. GI)

(A) K-shell (B) L-shell (C) M-shell (D) N-shell

20. Electronic configuration is based upon: (MLN. GI)

(A) Ionization energy (B) Electron affinity

(C) Mass number (D) Atomic number

21. Which one of the following is the cause for the discovery of proton? (SWL. GII)

(A) cathode rays (B) canal rays (C) x-rays (D) alpha rays

22. P sub-shell can accommodate electrons: (SGD. GII)

(A) 2 (B) 4 (C) 6 (D) 18

23. The atomic number of Argon (Ar) is: (RWP. GI)

(A) 16 (B) 10 (C) 8 (D) 18

Answers

1.	9.106×10^{-28}g	**2.**	Protons and neutrons	**3.**	Goldstein	**4.**	1.0087 amu
5.	Bohr	**6.**	Sir William Crooks	**7.**	Chadwick		
8.	Due to the ionization of gas molecule			**9.**	Rutherford	**10.**	negative
11.	neutron	**12.**	18	**13.**	2	**14.**	9
15.	M Shell	**16.**	32	**17.**	$1s^2, 2s^2, 2p^3$	**18.**	Argon
19.	N-shell	**20.**	Atomic number	**21.**	canal rays	**22.**	6
23.	18						

☆ **Give short answer to the following questions.**

1. Write down two properties of cathode rays. (LHR. GI, GRW. GI, MLN. GII)

Ans. Cathode rays are negatively charged. They travel in straight lines perpendicular to the cathode surface.

2. How does electron differ from neutron? (GRW. GII, BWP. GII)

Ans. Electron has negative charge, while neutron has no charge. Electron revolves around the nucleus, while neutron is in the nucleus. The mass of electron is 1840 times smaller than that of neutron.

3. State any two defects in Rutherford's atomic model. (GRW. GII, LHR. GII)

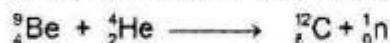
Ans. Following are the defects of this model.

i. According to classical theory, electrons being charge particles should release or emit energy continuously and they should fall ultimately into nucleus.

ii. If electrons emit energy continuously, they should form a continuous spectrum, but infact line spectrum was observed.

4. Who discovered neutron? Write its equation. (FBD. GI)

Ans. In 1932 Chadwick discovered neutron.

$${}^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + {}^{1}_{0}n$$

5. Write down the name of the particles which determine the mass of an atom. (FBD. G II, RWP. G II, 2014)

Ans. Neutron and Proton.

6. Write two properties of neutron particles. (SWL. GI, SGD. GII, RWP. GI)

Ans. i. It's mass is equal to proton mass.

ii. Neutron has no charge.

7. Who discovered proton and when? (GRW. GII)

Ans. In 1886, goldstein discovered proton

8. Write two differences between Rutherford's and Bohr's atomic theory.

Ans. Atomic theory of Rutherford: (FBD. GI & II)

i. It is based on classical theory.

ii. Electrons revolves around nucleus.

Atomic Theory of Neil Bohr:

i. It is based on quantum theory.

ii. Electron revolves around nucleus in specific orbits.

9. Find out the angular momentum of electron in the first orbit. (FBD. GI)

Ans. $mv = {}^{nh}/_{2\pi} = \dfrac{1 \times 6.63 \times 10^{-34}}{2 \times 3.14}$

$$= \frac{6.63 \times 10^{-34}}{6.28}$$

10. Why positive rays are called canal rays. (SGD. GI)

Ans. Goldstein observed the rays other than cathode rays in discharge tube. He observed that these rays move opposite to cathode rays. He used perforated cathode in discharge tube. These rays passed through this perforated cathode and produce

light on tube wall. He give them as "canal rays".

11. Write down any four properties of canal rays. (FBD. GI, RWP. GII)

Ans. i. These rays travel in straight line opposite to cathode rays.

ii. Their deflection in magnetic field shows that they have positive charge.

iii. The nature of canal rays depend upon the nature of gas present in discharge tube.

iv. The mass of these particles was found to be equal to proton or multiply of it.

12. How positive rays are generated? (DGK. GI)

Ans. When cathode rays colloids with remaining gas molecules in discharge tube, the gas molecules converted into positive ions.

$$M + e^- \longrightarrow M^- + 2e^-$$

13. What does Quantum mean? (DGK. GI)

Ans. Quantum means specific energy. It is small amount of energy that can absorb or release in the form of electromagnetic radiation.

14. What is plum pudding theory and who presented it? (RWP. GII, DGK. GII)

Ans. Thomson present plum pudding theory ,according to him atoms are solid structures of positive charge, with tiny negative particles stuck inside. It is like plum in pudding.

15. Write down the electronic configuration of Be and Ne. (LHR. GI)

Ans. $Be = 1s^2, 2s^2$;

$Ne = 1s^2, 2s^2, 2p^6$

16. Write the electronic configuration of "S". (GRW. GII, SWL. GII, DGK. GII)

Ans. $^{16}S = 1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

17. How many sub shells are there in second shell? (FBD. GII)

Ans. There are two sub shells "s" and "p" in second shell.

18. Differentiate between Shell and Subshell. (MLN. GI, FBD. GII, DGK. GII)

Ans. **Shell:** Electron revolves around the nucleus, according to their energy level, at different distances. These are called shells. These are represented by English letters K,L,M,N.

sub-shell: The points of shell, where there is more chance of occurrence of electron are called sub-shell or orbitals. These are represented by s,p,d and f.

19. Write down the electronic configuration of (i) Na (ii) Al (SGD.G I, SWL. GI)

Ans. $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

$_{13}Al = 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

20. Write down the electronic configuration of chlorine ($C\ell$). (RWP. GI, LHR. GII)

Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

21. Write down the electronic configuration of Nitrogen and Oxygen. (RWP. GII)

Ans. $_7N = 1s^2, 2s^2, 2p^3$

$_8O = 1s^2, 2s^2, 2p^4$

22. How many electrons will be in M shell of an atom having atomic number 15. (DGK. GII)

Ans. There will be 5 electrons.

23. Define electronic configuration. (BWP. GII)

Ans. The arrangement of electron around the nucleus is called electronic configuration.

24. Write the electronic configuration of Phosphorus ion $^{31}_{15}P^3$ and how many neutrons are in it. (GRW. GI, LHR. GII, RWP. GII)

Ans. $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

It has 16 neutrons.

25. Write down the number and names of sub-shells in N-shell. (GRW. GI)

Ans. The number of sub shell of N is 4, s, p, d, and f.

26. Write down the electronic configuration of nitrogen. (GRW. GI)

Ans. The electronic configuration of Nitrogen is $1s^2, 2s^2, 2p^3$.

27. Write electronic configuration of Mg^{2+} and $A\ell^{3+}$. (MLN. GI, GRW. GII)

Ans. $Al^{+3} = 1s^2, 2s^2, 2p^6$

$Mg^{+2} = 1s^2, 2p^2, 2p^6$

28. Write the Electronic Configuration of Oxygen. (MLN. GI)

Ans. $_8O = 1s^2, 2s^2, 2p^4$

29. Write the electronic configuration of silicon (Si) and Aluminium (Al) atoms. (MLN. GII)

Ans. Silicon: $_{14}Si = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$ **Aluminium:** $_{13}A\ell = 1s^2, 2s^2, 2p^6, 3s^2, 3p^1$

30. How many maximum electrons can be accommodated in L and M shells? (MLN. GII)

Ans. L shell has (8) eight, while M shell has eighteen (18) electrons.

31. An element has 5 electrons in M shell. Find out its atomic number. (SGD. GI)

Ans. It's atomic number is 15.

32. Write the electronic configuration of an element having 11 electrons. (BWP. GII)

Ans. $1s^2, 2s^2, 2p^6, 3s^1$

33. Write down the electronic configuration of Be and Ne. (BWP. GI)

Ans. Barellium $= 1s^2, 2s^2$

Neon $= 1s^2, 2s^2, 2p^6$

2.3 Isotopes

☆ Tick the correct answer.

1. **Which radioisotope is used for the diagnosis of goiter in thyroid gland?** (FBD. GI)
   (A) Cobalt-60 (B) Iodine-131 (C) Stronitium-90 (D) Phosphorus-30
2. **When U.235 breaks up, it produces.** (SWL. GI)
   (A) electrons (B) neutrons (C) protons (D) none of these
3. **Deutrium is used to make:** (RWP. GI & II 2014)
   (A) Light water (B) Heavy water (C) Soft water (D) Hard water
4. **How many stable isotopes are present in carbon:** (DGK GII)
   (A) one (B) two (C) three (D) four

Answers

1. Iodine-131 2. neutrons 3. Heavy water 4. two

☆ Give short answer to the following questions.

1. **Define the term carbon dating.** (LHR. GI, SGD. GI, RWP. GI, LHR. GI, SWL. GII)

Ans. The important method of age determination of old carbon containing objects (fossils) is called carbon dating.
It depends on measurement of radioactivity of C-14 present in the fossils.

2. **State any two uses of isotope.** (LHR. GII, RWP. GII, SWL. GI, RWP. GI)

Ans. **Radiotherapy:** Different isotopes like Sr-90 and P-32 of elements are used for treatment of skin cancer.
Use in power generation: Radioactive isotopes are used for production of electricity in nuclear reactor through controlled nuclear fission reaction.

3. **For what purpose U-235 is used?** (LHR. GII, LHR. GII, MLN. GI)

Ans. In nuclear reactor, slow moving neutrons are bombarded on uranium to generate electricity, for this purpose nuclear fission reaction is used.

4. **Complete the chemical equation.** (GRW. GI)

$$^{9}_{4}Be + ^{4}_{2}He \longrightarrow ______$$

Ans. $^{9}_{4}Be + ^{4}_{2}He \longrightarrow ^{12}_{6}C + ^{1}_{0}n$

5. **Complete the chemical equation.** $^{235}_{92}U + ^{1}_{0}n \longrightarrow ? + ? + ?$ (MLN. GI, FBD. GI)

Ans. $^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n$ + energy.

6. **Write down the names of two Isotopes of Chlorine.** (MLN. GII)

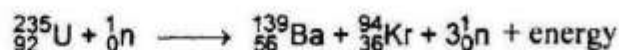
Ans. The two isotopes of chlorine are ^{35}Cl and ^{37}Cl.

7. **What happens when slow moving neutrons hit the Uranium. Write chemical equation.** (SGD.GI)

Ans. $^{235}_{92}U + ^{1}_{0}n \rightarrow ^{139}_{56}Ba + ^{94}_{36}K + 3^{1}_{0}n$

8. What is Nuclear fission reaction? Give example. (DGK. GI, SGD. GII)

Ans. When slow moving neutrons are bombarded on nucleus, it is divided into two small nuclei with the emission of energy. This process is known as nuclear fission reaction.

$$^{235}_{92}U + ^{1}_{0}n \longrightarrow ^{139}_{56}Ba + ^{94}_{36}Kr + 3^{1}_{0}n + \text{energy}$$

9. Write the use of I-131. (DGK. GI, SWL. GII)

Ans. To diagnose goiter, the isotope of Iodine (131) is used as a tracer.

10. Write down the chemical equation for the bombardment of α-particles on beryllium target. (DGK. G II, 2014)

Ans. $^{9}_{4}Be + ^{4}_{2}He \rightarrow ^{12}_{6}C + ^{1}_{0}n$

11. A patient has Goiter. How will it be detected? (BWP. GI)

Ans. To diagnose the goiter in thyroid gland, isotope of iodine (I - 131) is used as a tracer.

12. Define Isotopes. Give two examples. (FBD. GI, GRW. GI, MLN. GI)

Ans. The atoms of element having same atomic number but different mass number are known as isotopes. e.g

(i) The isotopes of carbon are $^{12}C, ^{13}C, ^{14}C$. (ii) Isotopes of chlorine ^{35}Cl, ^{37}Cl

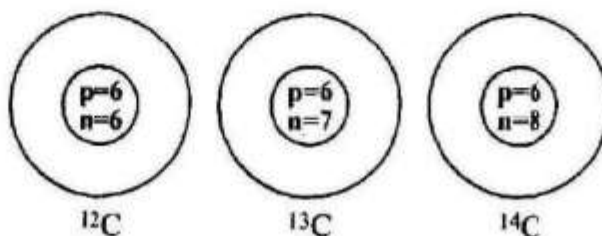
13. Describe the uses of isotopes in radio therapy. (SGD. GII)

Ans. i. For the treatment of skin cancer, different type of isotopes P-32, Sr-90 are used, because these emit less penetrating β-radiation.

ii. Co-60 is used for cancer treatment, because it emit more penetrating γ(gamma) radiations.

14. Explain the isotopes of carbon. (RWP. GI)

Ans. The two isotopes of carbon ^{12}C and ^{13}C are stable, while ^{14}C is radioactive. Naturally, the amount of ^{12}C is 98.9%, while the amount of both ^{13}C and ^{14}C is 1.1%. They all have same numbr of electron and protons but different number of neutrons.

^{12}C ^{13}C ^{14}C

PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Major Concepts:

3.1 Periodic Table
3.2 Periodic Properties

Time allocation

Teaching periods	12
Assessment periods	02
Weightage	10%

Students Learning Outcomes:

Students will be able to:

- Distinguish between period and group in the Periodic table.
- State the Periodic law.
- Classify elements (into two categories: groups and periods) according to the configuration of their outer most electrons.
- Determine the demarcation of the periodic table into s-block and p-block.
- Explain the shape of the periodic table.
- Determine the location of families of the periodic table.
- Recognize the similarity in the physical and chemical properties of elements in the same family of the elements.
- Identify the relationship between electronic configuration and position of elements in the periodic table.
- Explain how shielding effect influences periodic trends.
- Describe how electronegativities change within a group and within a period in the periodic table.

3.1 PERIODIC TABLE

Q.1 Explain the following.

(i) Dobereiner's Law of Triads ***(ii) Newlans Law of Octave***

(iii) Mendelee's Periodic Table

Ans. Dobereiner's Law of Triad: A Germen chemist Dobereiner arranged chemically similar element in group of three on the basis of their atomic masses. These groups were called triads.

According to this law: "In a triad of similar elements, the atomic mass of the middle element is approximately the avere ge of the atomic masses of the other two elements. This is known as a law of triad.

Examples of Triads

Element	Atomic mass	Average atomic mass of Ist and 3rd elements
1 *Li*	7	
Na	23	$\frac{7+39}{2}=23$
K	39	
2 *Ca*	40	
Sr	88	$\frac{40+137}{2}=88.5$
Ba	137	

Defects: Since only few elements could be arranged in such groups (triads) hence this classification did not get wide acceptance.

(ii) Newlands Law of Octaves: In 1864 British chemist Newlands discovered a relationship between the atomic masses and the properties of elements and reported his law of octave.

Statement: If the elements are arranged in order of their increasing atomic masses, the properties of every 8th element, starting from any point are similar to that of first, this arrangement was named as Newlands law of octaves. He compared it with musical notes.

Defects:

Newlands law of octaves did not get much recognition as no space was left for undiscovered elements. The noble gases were also not known at that time.

(iii) Mendeleev's Periodic Table: This law states that the properties of the elements are periodic functions of their atomic masses.

Mendeleev (1834 - 1907) was a Russian chemist and inventor. He was the creator of first version of periodic table of elements. With help of the talle, he predicted the properties of elements yet to be discovered.

Explanation: This law was put forward by a Russian chemist, Mendeleev.

He arranged the known elements (only 63) in order of increasing atomic masses in horizontal rows called periods. The elements with similar properties were placed in same vertical columns. This arrangement of elements was called periodic table.

Defects: It can not explain the position of isotopes.

The wrong order of the atomic masses of some elements suggested that atomic

mass of an element cannot serve as the basis for the arrangement of elements.

Do you know?
Atomic number is a more fundamental property than atomic mass because atomic number of every element is fixed and it increases regularly by 1 from element to element. No two elements can have the same atomic number.

Test yourself 3.1:

i. What was the contribution of Dobereiner towards classification of elements?

Ans. Dobereiner classified the known elements into group of three similar elements called triads. According to him when elements are arranged in increasing order of their atomic masses, then the atomic mass of the middle element is approximately equal to the average of the other two elements of triads.

ii. How Newlands arranged the elements?

Ans. According to Newland, when the elements are arranged in increasing order of their atomic masses then the properties of every eighth element are similar to that of first one starting from any point.

iii. Who introduced the name Periodic Table?

Ans. Mendeleev introduced the name periodic table.

iv. Why the improvement in Mendeleev's periodic table was made?

Ans. The improvement in Mendeleev's periodic table was made due to following reasons.
(i) Mendeleev Periodic table does not explain the position of isotopes.
(ii) When the elements are arranged in increasing order of their atomic masses. Then order of certain elements become reversed.

v. State Mendeleev's periodic law.

Ans. Mendeleev's Periodic law: When the elements are arranged in increasing order of their atomic masses, then the properties of the elements are periodic function of their atomic masses.

vi. Why and how elements are arranged in a period?

Ans. The elements are arranged in a period on the basis of their atomic numbers and electronic configuration to make the study of elements and their compounds easier.

Q.2 Explain followings.

Ans. (i) Periodic Law (ii) Modern Periodic Table

(i) Periodic Law: "The properties of the elements are periodic functions of their atomic numbers".

Explanation: In 1913 H. Moseley discovered a new property of the elements i.e. atomic number.

He observed that atomic number instead of atomic mass should determine the position of element in the period table. The atomic number provides the basis of electronic configuration.

(ii) Modern Periodic Table: The modern periodic table is based on the increasing order of atomic number. The electronic configuration of atoms played an important role in the arrangement of periodic table.

The present form of the periodic table is called long form of the periodic table because it contains eighteen group.

When the elements are arranged according to increasing atomic number from left

to right in a horizontal row, properties of elements were found repeating after regular intervals such that elements of similar properties and similar configuration are placed in same group.

Q.3. Write down the main (salient) features of long form of periodic table.

Ans. (i) This table consists of seven horizontal rows called periods.

(ii) First period consists of only two elements. Second and third periods consist of 8 elements each. Fourth and fifth periods consist of 18 elements each. Sixth period has 32 elements while seventh period has 23 elements and is incomplete.

(iii) Elements of a period show different properties.

(iv) There are 18 vertical columns in the periodic table numbered 1 to 18 from left to right, which are called groups.

(v) The elements of a group show similar chemical properties.

(vi) Elements are classified into four blocks depending upon the type of the sub-shell which gets the last electron.

Modern Periodic Table or long form of the Periodic Table of Elements

Q.4. Define the different blocks in the modern periodic table.

Ans. Blocks of elements in the periodic table: The elements in the periodic table are divided into four blocks known as s, p, d and f-blocks.

This classification of elements in based upon the electronic configuration.

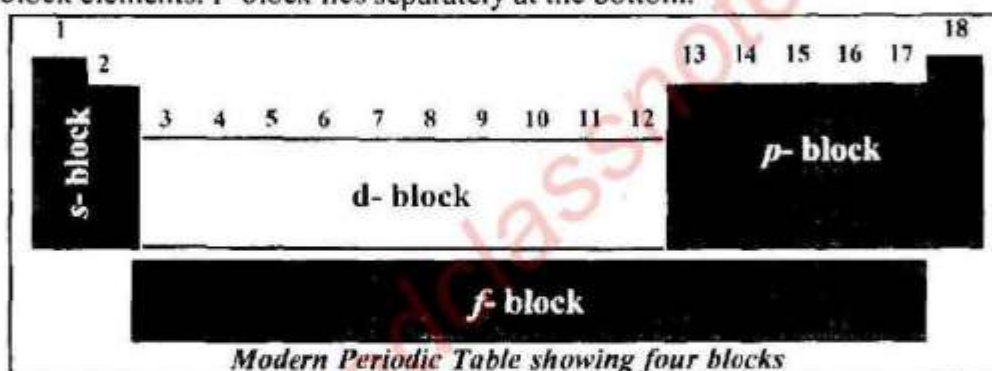
The type of the sub shell which receives the last electron of an atom determines the block to which that element belongs.

1. **s-block elements:** The elements whose valence electrons enter into the s-orbital are called s-block elements group IA and group IIA constitute s-block.

2. **p-block elements:** The elements whose valence electrons enter in p-orbitals are called p-block elements. Group III- A to VII - A and zero group [except He] constitute p-block.

3. **d-block elements:** The elements whose valence electrons enter into d-orbitals are called d-block elements. Group III-B to VIII-B and I-B and II-B constitute d-block.

4. **f-block elements:** The elements whose valence electrons enter into f-orbital are called f-block elements. F-block lies separately at the bottom.

Modern Periodic Table showing four blocks

Do you know?
Alchemy! For thousand years alchemy years alchemy remained field of interest for the scientists. They worked with two main objectives; change common metals into gold and second find cure to diseases and give eternal life to people. They believed all kinds of matter were same combination of four basic elements. Substances are different because these elements combine differently. Changing composition or ratio of any one element, new substances can be formed. The way of making gold from silver or lead was never found and secret of eternal life was never discovered. However, many methods and techniques invented by alchemists are still used in chemistry.

Q.5. What is meant by periods? Explain the periods of periodic table.

Ans. Periods: Horizontal rows of the elements in the periodic table is called period.

Number of periods in modern period table: There are seven periods in the modern periodic table.

Explanation:

1. **First Period:** First period is called short period. It contains two elements hydrogen and helium.

2. **Second Peroid:** Second period is called normal period. It contains eight elements. Lithium, Beryllium, Boron, Carbon, Nitrogen, Oxygen, Fluorine and Neon.

3. **Third period:** Third period is also called normal period. It contains eight elements.

4. **Fourth period:** It is called long period. It consists of eighteen elements.

5. **Fifth period:** It is also called long period. it consists of eighteen elements.

6. **Sixth period:** Sixth period is called very long period. It consists of 32 elements. [It contains 8 normal, 10 outer transition and 14 inner transition elements]. The series of the fourteen elements [La^{57} to Lu^{71}] in sixth period is called Lanthanides.

7. **Seventh period:** Seventh period is also a very long period.
It contains a series of fourteen element [Ac^{89} to Lr^{103}] called actinides.

Period No.	Name of the Period	Number of Elements	Range of Atomic Numbers
1st	Short Period	2	1 to 2
2nd	Normal Period	8	3 to 10
3rd		8	11 to 18
4th	Long Period	18	19 to 36
5th		18	37 to 54
6th	Very Long Period	32	55 to 86
7th		[23]*	87 to 118*

*Since new elements are expected to be discovered, it is an incomplete period

All the periods except the first period start, with an alkali metal and even at a noble gas.

It is observed that number of elements in a period is fixed because of maximum number of electrons which can be accommodated in the particular valence shell of the elements.

Q.6(a) What is meant by groups explain the groups?

(b) What is meant by transition element? [Define their types]

Ans. Group: The vertical columns in the periodic table are called groups.

Elements with similar outer electronic configuration show similar properties and are placed in one group.

Different Groups of the Periodic Table

Valence electrons	Group number	Family name	General Electronic configuration
1 electron	1	Alkali metals	ns^1
2 electron	2	Alkaline earth metals	ns^2
3 electron	13	Boron family	$ns^2\ np^1$
4 electron	14	Carbon family	$ns^2\ np^2$
5 electron	15	Nitrogen family	$ns^2\ np^3$
6 electron	16	Oxygen family	$ns^2\ np^4$
7 electron	17	Halogen family	$ns^2\ np^5$
8 electron	18	Noble gases	$ns^2\ np^6$

Group I: It consists of

Hydrogen, Lithium, Sodium, Potassium, Rubidium, Cesium, Francium.

All the elements of group I have similar electronic configuration in their last shell hence they are called family. This family is known as alkali metals (except hydrogen).

The groups 1 and 2 and 13 to 17 contain the normal elements.

The groups 3 to 12 are called transition elements.

(b) Transition elements: The elements in which "d" or "f" orbitals are under the process of completion are called transition elements e.g. iron.

Types of transition elements: They are divided into two groups.

1. Outer transition elements: The elements in which d-orbital are under the process of completion are known as outer transition elements. They are also known as d-block elements. They are placed at the center of the periodic table.

2. Inner transition elements: The elements in which f-orbitals are under the process of completion. They are also known as f-block elements. They are placed at the bottom of the periodic table.

Do you know?

Beautiful fireworks display are common on celebrations like Pakistan Day or even on marriages. A technology invented in China is used all over the world. It is dangerous but careful use of various elements and particularly metal salts of different composition give beauty and colors to the fireworks. Elements like magnesium, aluminium are used in powdered form. Salts of sodium give yellow color, calcium - red; strontium-scarlet; barium-green and copper-bluish green. Usually nitrates and chlorates are used. Other chemicals are added to give brilliance and different shades. Because of fire hazard and risk to life and property, only skilled professionals use them.

Test yourself 3.2:

i. How the properties of elements repeat after regular intervals?

Ans. The properties of the elements depends upon their electronic configuration. The properties of the elements repeated after a regular interval due to the repeating trend in their electronic configurations.

ii. In which pattern modern periodic table was arranged?

Ans. The modern periodic table was arranged on the basis of atomic numbers and electronic configuration of the elements.

iii. How many elements are in first period and what are their names and symbols?

Ans. There are two elements in the first period.

(i) Hydrogen (H) (ii) Helium (He)

iv. How many elements are placed in 4th period?

Ans. There are eighteen elements on the 4th period.

v. From which element lanthanide series starts?

Ans. Lanthanide series start from Lanthanum (La - 57).

vi. From which period actinides series starts?

Ans. The actinides series start from 7th period.

vii. How many elements are in 3rd period, write their names and symbols?

Ans. There are total eight elements in the third period.

Name of element	Symbols	Name of element	Symbols
Sodium	Na	Magnesium	Mg
Aluminium	Al	Silicon	Si
Phosphorous	P	Sulphur	S
Chlorine	Cl	Argon	Ar

viii. **How many periods are considered normal periods?**

Ans. Two periods [2nd and 3rd] are known as normal periods.

ix. **What do you mean by a group in a periodic table?**

Ans. **Group:** The vertical arrangement of elements in the periodic table is called group.

x. **What is the reason of arranging elements in a group?**

Ans. The elements are arranged in a group in order to make the study of elements and their compounds easier.

xi. **What do you mean by periodic function?**

Ans. The properties of the elements of which are repeated again and again after a regular interval are called periodic properties or periodic functions.

xii. **Why the elements are called s or p block elements?**

Ans. **s-block elements:** The elements whose valence electrons enter into the "s" orbitals are called s-block elements e.g. Helium.

p-block elements: The elements whose valence electrons enter into the p-orbital are called p-block elements e.g. oxygen.

xiii. **Write down the names of elements of group 1 with their symbols?**

Ans.

Name of element	Symbols	Name of element	Symbols
Hydrogen	H	Lithium	Li
Sodium	Na	Potassium	K
Rubidium	Rb	Cesium	Cs
Francium	Fr		

xiv. **How many members are in group 17, is there any liquid, what is its name?**

Ans. There are total five members in group 17.

Liquid element: Bromine (Br)

3.2 PERIODICITY OF PROPERTIES

Q.7. What is meant by periodicity of properties? What is meant by atomic radius (atomic sizes)? Describe its trends along periodic table.

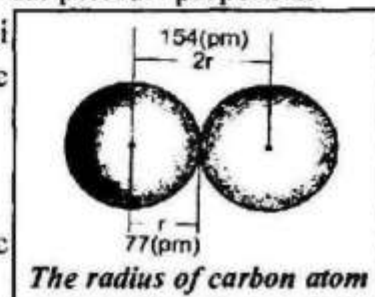
Ans. Periodicity: The physical and chemical properties of elements that are repeated after a regular interval are called periodic properties and the phenomenon is called periodicity of properties e.g. atomic radius, ionization energy are periodic properties.

Atomic Radius: The half of the distance between the nuclei of two similar adjacent (bonded) atoms is called atomic radius.

OR

The average distance between the nucleus of an atom and its outermost electronic shell is called atomic radius.

Units: Very small units of length are used to measure atomic radius e.g. picometer (pm).

The radius of carbon atom

Trends along periodic table

1. **Trends along group:** The atomic radius (atomic size) increases from top to bottom in a group due to an addition of one more electronic shell at each step down the group which decreases the effective nuclear charge.

Table

1st group elements	Atomic radii (pm)
^{3}Li	152
^{11}Na	186
^{19}K	227
^{37}Rb	248
^{55}Cs	265

2. **Trends along period:** Atomic radius decreases from left to right in a period due to increase in nuclear force and decreases in atomic size.

The effective nuclear charge gradually increases from left to right in a period which pulls down the outermost (last) shell towards the nucleus.

Table

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Atomic radii (pm)	152	113	88	77	75	73	71	69

Trends of atomic size of transition elements (atomic radius):

The trends of atomic size of transition elements has slight variation when we consider the transition elements.

The atomic radius (atomic size) of the transition elements first reduces and then there is an increase in it when we move from left to right in 4th period.

Q.8.(a) What is meant by shielding effect? Describe its trends along periodic table.

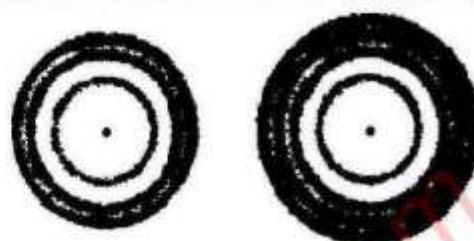
(b) What is meant by ionization energy? Describe its trends along periodic table.

Ans. Shielding effect or screening effect:

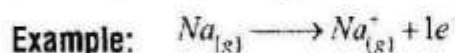
The decrease in the attractive force exerted by the nucleus on the valence electrons due to the presence of the electrons lying between the nucleus and valence-shell is called shielding effect.

Trends along periodic Table:

Trends along group: Shielding effect increases from top to bottom in a group due to increase in the number of inner-shells or inner-shell electrons.

Trends along periods: The shielding effect does not change in a period. This is because from left to right in a period, the number of inner shells remain the same.

Shielding effect is more in potassium atom than that of sodium atom.

(b) Ionization energy: The amount of energy required to remove the outermost electron from an isolated gaseous atom in its ground state is called ionization energy (First ionization energy)

Example: $Na_{(g)} \longrightarrow Na^{+}_{(g)} + 1e^{-}$

I. E = + 496 kJ/mole

Second ionization energy: The energy required to remove second electron from a positive ion is called second ionization energy.

Units: The ionization energy is expressed in kJ/ mole.

Trend along Periodic Table:

Trends along group: The ionization energy decreases down the group due to increase in atomic size.

1st group elements	Ionization energy ($kJmol^{-1}$)
^{3}Li	520
^{11}Na	496
^{19}K	419
^{37}Rb	403
^{55}Cs	377

Trends along period: The ionization energy increases from left to right in a period due to increases in nuclear force and decrease in atomic size.

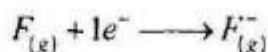
2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Ionization energy ($kJmol^{-1}$)	520	899	801	1086	1402	1314	1681	2081

Q.9.(a) Define electron affinity and give its variation in groups and periods.
(b) What is meant by electronegativity? Describe its trends along periodic table.

Ans. Electron affinity: The maximum amount of energy which is released or absorbed when an electron is added to an isolated gaseous atom to form a negatively charged ion is

called the electron affinity.

Example:

$$F_{(g)} + 1e^- \longrightarrow F^-_{(g)}$$

$\Delta H = -328$kJ/mole

Unit: Electron affinity is measured in kJ/mole.

Variation in periodic Table:

Trends along group: Electron affinity gradually decreases down the group due to increase in atomic size.

Shielding effect increases down the group which decreases the attraction for incoming electrons hence less energy is released.

17th group elements	Electron affinity ($kJmol^{-1}$)
^{9}F	-328
^{17}Cl	-349
^{35}Br	-325
^{53}I	-295

Trends along period: Electron affinity increases from left to right in a period due to decrease in atomic size.

When the size of the atom decreases, the attraction for incoming electrons increases hence more energy is released.

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Electron affinity ($kJmol^{-1}$)	-60	>0	-29	-122	0	-141	-328	0

(b) Electronegativity: The ability of an atom to attract the shared pair of electrons towards itself is called electronegativity.

Trends along periodic table:

Trends along group: Electronegativity decreases down the group due to increase in atomic size which decreases the attraction for the shared pair of electrons.

17th group elements	Electronegativity
^{9}F	4.0
^{17}Cl	3.2
^{35}Br	3.0
^{53}I	2.7

Trends along period: The electronegativity increases from left to right in a period due to decrease in atomic size which increases the attraction for shared pair of electrons.

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F
Electro negativity	1.0	1.6	2.0	2.6	3.0	3.4	4.0

Test yourself 3.3

i. Define atomic radius?

Ans. **Atomic radius:** Half of the distance between the nuclei of two similar adjacent atoms is called atomic radius.

ii. What are SI units of atomic radius?

Ans. Pico meter (pm)

iii. Why the size of atoms decreases in a period?

Ans. The size of the atoms decreases down in a period due to increase in nuclear force.

iv. Define ionization energy.

Ans. **Ionization Energy:** The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom in its ground state is called ionization energy.

v. Why the 2nd ionization energy of an element is higher than first one?

Ans. The 2nd ionization energy is greater than first because when an electron is removed, the number of electrons is decreased while the number of protons remain same. So the size of atom decreases which increased the ionization energy.

vi. What is the trend of ionization energy in a group?

Ans. The ionization energy decreases down the group due to increase in atomic size.

vii. Why the ionization energy of sodium is less than that of magnesium?

Ans. Ionization energy of sodium is less than that of magnesium because magnesium has greater nuclear charge than sodium.

viii. Why is it difficult to remove an electron from halogens?

Ans. It is difficult to remove an electron from halogens due to their very high values of ionization energies.

ix. What is shielding effect?

Ans. **Shielding effect:** The decrease in attractive force exerted by the nucleus on the valence shell electrons due to the presence of electrons lying between the nucleus and the valence shell.

x. How does shielding effect decrease the forces of electrostatic attractions between nucleus and outer most electrons?

Ans. The electrons present between the nucleus and the outermost shell of an atom reduces the nuclear charge felt by the electrons present in the outermost shell. The attractions of outer electrons towards nucleus is partially reduced because of presence of inner electrons. As a result an atom experiences less nuclear charge than that of actual charge.

xi. Why does the bigger size atoms have more shielding effect?

Ans. Bigger size atoms have more shielding effect due to the presence of more shells between the nucleus and valence shell.

xii. Why does the trend of electron affinity and electronegativity is same in a period?

Ans. Electron affinity and electronegativity have same trend in a period because both depends upon the atomic size and nuclear charge. When we move from left to right, atomic size decreases and nuclear charge increases which increases the electron affinity and electronegativity.

xiii. Which element has the highest electronegativity?

Ans. Fluorine has the highest value of electronegativity.

Key Points

- In nineteenth century attempts were made to arrange elements in a systematic manner.
- Dobereiner arranged elements in group of three called triads.
- Newlands arranged elements in groups of eight like musical notes.
- Mendeleev constructed Periodic Table containing periods and columns, by arranging elements in order of increasing atomic weights.
- There are total eighteen groups and seven periods in the modern Periodic Table.
- Depending on outermost electrons and electronic configuration, element in periodic table are grouped in s, p, d and f blocks.
- Atomic size increases down a group but decreases along the period.
- Shielding effect is greater in atoms with greater number of electrons.
- Electronegativity increases a long a period and decreases down the group.

Exercise (Solved)

Multiple Choice Questions

Put a (✓) on the correct answer.

1. **The atomic radii of the elements in Periodic Table:**
   (a) increase from left to right in a period (b) increase from top to bottom in a group
   (c) do not change from left to right in a period
   (d) decrease from top to bottom in a group
2. **The amount of energy given out when an electron is added to an atom is called:**
   (a) lattice energy (b) ionization energy (c) electronegativity(d)electron affinity
3. **Mendeleev Periodic Table was based upon the:**
   (a) electronic configuration (b) atomic mass
   (c) atomic number (d) completion of a subshell
4. **Long form of Periodic Table is constructed on the basis of:**
   (a) Mendeleev Postulate (b) atomic number (c) atomic mass (d) mass number
5. **4th and 5th period of the long form of Periodic Table are called:**
   (a) short periods (b) normal periods (c) long periods (d)very long periods
6. **Which one of the following halogen has lowest electronegativity?**
   (a) flourine (b) chlorine (c) bromine (d) iodine
7. **Along the period, which one of the following decreases:**
   (a) atomic radius (b)ionization energy (c) electron affinity (d) electronegativity
8. **Transition elements are:**
   (a) all gases (b) all metals (c) all non-metals (d) all metalloids

9. **Mark the incorrect statement about ionization energy:**
(a) it is measured in $kJmol^{-1}$ (b) it is absorption of energy
(c) it decreases in a period (d) it decreases in a group

10. **Point out the incorrect statement about electron affinity:**
(a) it is measured in $kJmol^{-1}$ (b) it involves release of energy
(c) it decreases in a period (d) it decreases in a group

Answers: **1.** increase from top to bottom **2.** electron affinity **3.** atomic mass
4. atomic number **5.** long periods **6.** iodine **7.** atomic radius
8. all metals **9.** It decreases i a period **10.** It decreases in a period

Short Answer Questions.

1. **Why are noble gases not reactive?**

Ans. The noble gases are not reactive because they have completely filled valence shells and do not react with other elements to form compounds.

2. **Why Cesium (at. no. 55) requires little energy to release its one electron present in the outermost shell?**

Ans. The cesium required little energy to release its last electron due to greater atomic size and larger shielding effect.

3. **How is periodicity of properties dependent upon number of protons in an atom?**

Ans. For answer see Q. 7.

4. **Why shielding effect of electrons makes cation formation easy?**

Ans. The shielding effect of electrons make cation [positive ion] formation easy because when the shielding effect increases it decreases the attractive force between valence electron and nucleus, which make the removal of electron easy and atom easily converted into positive ion [cation].

5. **What is the difference between Mendeleev's periodic law and modern periodic law?**

Ans.

Mendeleev's Periodic Law	Modern Periodic law
1. Mendeleev's periodic law is based upon the atomic mass.	1. Modern periodic law is based upon atomic number.
2. Mendeleev's periodic law does not show the position of isotopes.	2. There is no need for separate position of isotopes in the modern periodic law.

6. **What do you mean by groups and periods in the Periodic Table?**

Ans. **Groups:** The vertical arrangement of elements in the periodic table is called group.
Periods: The horizontal arrangement of elements in the periodic table are called periods.

7. **Why and how are elements arranged in 4th period?**

Ans. For answer see Q.5

8. **Why the size of atom does not decrease regularly in a period?**

Ans. The size of atom does not increase regularly in these periods which contain

transition elements due to the presence of d or f orbital which shows poor shielding effect.

9. Give the trend of ionization energy in a period.

Ans. For answer see Q. 8(b) [Trend along period only]

Long Answer Questions

Q.1 Explain the contribution of Mendeleev for the arrangement of elements in his Periodic Table.

Ans. For answer see Q. 1

Q.2 Show why in a 'period' the size of an atom decreases if one moves from left to right.

Ans. For answer see Q. 7

Q.3 Describe the trends of electronegativity in a period and in a group.

Ans. For answer see Q. 9(b)

Q.4 Discuss the important features of modern Periodic Table.

Ans. For answer see Q. 3

Q.5 What do you mean by blocks in a periodic table and why elements were placed in blocks?

Ans. For answer see Q. 4

Q.6 Discuss in detail the periods in Periodic Table?

Ans. For answer see Q. 5

Q.7 Why and how elements are arranged in a Periodic Table?

Ans. Before nineteenth century only few elements were discovered and they can be studied individually. With the passage of time new elements were discovered. More and more new compounds were also prepared. Now it became difficult to study elements and their compounds. To solve this problem it is essential to classify the elements.

The elements are classified into groups and periods according to their properties. By classification, the study of all elements and their compounds become easier. Many chemists contributed for classification of elements [from Dobereiner to Mendeleev and Mosely].

Q.8 What is ionization energy? Describe its trend in the Periodic Table?

Ans. For answer see Q. 8(b)

Q.9 Define electron affinity, why does it increase in a period and decrease in a group in the Periodic Table.

Ans. For answer see Q. 9(a)

Q.10 Justify the statement, bigger size atoms have more shielding effect thus low ionization energy.

Ans. For answer see Q. 8

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

3.1 Periodic Table

☆ **Tick the correct answer.**

1. **How many blocks are in modern periodic table:** (LHR. GII, FBD. GI)
(A) 3 (B) 4 (C) 5 (D) 6
2. **The vertical columns in the periodic table are called:** (GRW. GI)
(A) periods (B) atomic number (C) groups (D) atomic mass
3. **The number of groups in periodic table is:** (GRW. GII)
(A) 8 (B) 9 (C) 18 (D) 12
4. **Modern periodic law is presented by:** (FBD. GI)
(A) Dobereiner (B) Newlands (C) Mendeleev (D) H.Moseley
5. **Group 17th belongs:** (FBD. GI)
(A) Halogens (B) Noble gases (C) Alkali metals (D) None of these
6. **1st period has elements:** (FBD. GII, DGK. GI)
(A) 2 (B) 3 (C) 4 (D) 5
7. **Long form of periodic Table is constructed on the basis of:** (MLN. GI, SWL.GII)
(A) Mendeleev Postulate (B) Atomic Number
(C) Atomic Mass (D) Mass Number
8. **Second period contains number of elements:** (SWL. GII)
(A) 2 (B) 8 (C) 18 (D) 32
9. **Transition elements are:** (SGD. GI, DGK. GI & II)
(A) All gases (B) All Metals (C) All metalloids (D) All non metals
10. **The first period consists of:** (SGD. GII)
(A) Two elements (B) Three elements (C) Four elements (D) Five elements
11. **4th and 5th periods of long form of periodic table are called:** (RWP. GI)
(A) Short periods (B) Normal periods (C) Long periods (D) Very long periods
12. **Forth Period contains number of elements:** (DGK. GI, BWP. GII, RWP. GII)
(A) 2 (B) 8 (C) 18 (D) 32
13. **The horizontal lines are called:** (DGK. GI, MLN. GII)
(A) Periods (B) Atomic number (C) Groups (D) Atomic mass
14. **Mendeleev periodic table was based on the:** (DGK. GII, RWP. GI)
(A) electronic configuration (B) atomic mass
(C) atomic number (D) completion of a subshell

15. Number of elements in normal period are: (LHR. GI & II, BWP. GII)
(A) 18 (B) 10 (C) 8 (D) 32

16. The number of elements is 6th period is: (GRW. GII, SWL. GI)
(A) 18 (B) 32 (C) 54 (D) 80

17. How many elements are present in third period? (FBD. GII)
(A) 2 (B) 4 (C) 8 (D) 18

18. The elements of group seventeen are called: (FBD. GII)
(A) Carbon family (B) Nobel gases (C) Alkaline earth metals (D) Halogens

19. The concept of triad was presented by: (MLN. GI)
(A) Dobereiner (B) Newlands (C) Mendeleev (D) Moseley

20. The elements of first group are known as: (SWL. GI)
(A) Alkali metals (B) Alkaline earth metals
(C) metalloids (D) halogens

21. Carbon family has general electronic configuration: (SGD. GII)
(A) ns^2np^1 (B) ns^2np^2 (C) ns^2np^3 (D) ns^2np^4

22. Transition metals are found in block: (RWP. GII)
(A) s (B) p (C) d (D) f

Answers

1. 4	2. groups	3. 18	4. H.Moseley
5. Halogens	6. 2	7. Atomic Number	
8. 8	9. All Metals	10. Two elements	11. Long periods
12. 18	13. Periods	14. atomic mass	15. 8
16. 32	17. 8	18. Halogens	19. Dobereiner
20. Alkali metals	21. ns^2np^2	22. d	

☆ **Give short answer to the following questions.**

1. Define Mendleev's Periodic Law. (LHR. GI, GRW. GII, LHR. GI, MLN. GI)

Ans. The law states that "the characteristics of elements are periodic functions of their atomic masses."

2. Define periods and groups in the periodic table. (LHR. GII, SWL. GII)

Ans. In periodic table the horizontal rows are called periods while vertical columns are called group.

3. Write names of any four elements of group 17. (GRW. GI, SWL. GI)

Ans. Chlorine, Bromine, Fluorine, Iodine.

4. Write down the name or symbol of the elements of first period. (FBD. GII)

Ans. In first period only two elements are present Hydrogen (H) and Helium (He).

5. Write down names of any four element in Group-1st. (MLN. GI)

Ans. Hydrogen, Lithium, Sodium, Potassium

6. Define Newland's Law of Octaves. (MLN. GII, RWP. GI, DGK. GII)

Ans. In 1864 an English chemist Newland gave law of octaves. According to him If elements are arranged in order of their increasing atomic masses, the properties of every 8th element, starting from any point are similar to one from.

7. Define period with one example. (SGD. GII)

Ans. The horizontal rows of elements in periodic table called periods. For example, those elements having one electron in their valence shell called alkali metals.

8. Define Mosely's Periodic Law. (DGK. GII)

Ans. He states that,
"Properties of elements are periodic function of their atomic numbers."

9. What is contribution of Doberiener towards classification of elements? (DGK. GII)

Ans. A German chemist Doberiener observed the relation between atomic masses of triads (group of three elements). In triad, the atomic mass of the middle element is approximately the average of atomic masses of the other two elements. For example group of triad consist of calcium, stronsium and Barium. The atomic mass of stronsium is the average of atomic masses of calcium and barium. Since few elements could be classified, so this type of classification did not get wide acceptance.

10. Differentiate between Period and Group. (BWP. GI & II)

Ans. The horizontal rows in periodic table called periods. While the vertical columns are known as groups.

11. Write down general electronic configuration of carbon family. (LHR. GI)

Ans. ns^2, np^2

12. What are the elements arranged in group 3 to 12 called? (LHR. GII)

Ans. From third (3rd) to twelvth (12th) groups element are known transition elements.

13. From which element lanthanide series starts? What is its atomic number? (GRW. GII)

Ans. Lanthanide series started with lanthanum atomic number of 57.

14. Write the names of elements arranged in group first of periodic table. (FBD. GI)

Ans. The first group of periodic table consist of Hydrogen, Lithium, Sodium, Potassium, Rubidium, Cesium and Francium.

15. How Newland arranged the elements? (MLN. GII)

Ans. Newland arranged elements, according to their increasing number of atomic masses. He observed that if elements arranged in octaves manner, the chemical properties of every 8th element resembles the first one.
His law did not get wide acceptance.

16. Write the names of any four P-block elements. (SWL. GII)

Ans. Boron, Carbon, Nitrogen, Oxygen.

17. Give reason that elements of group 13 to 18 are called p-block elements. (SGD. GI)

Ans. The elements from group 13-18 are known as p-block element, because, their valence electrons are found in p-sub shell.

18. Write down demerits of mendeleev's periodic table. (SGD. GII)

Ans. 1. Mendleeve's periodic table does not explain the position of isotopes.
2. And when the elements are arranged according to their increase atomic masses, the order of certain elements become reversed.

19. Define periodic law and modern periodic table. (RWP. GI, SGD. GII)

Ans. **Periodic law:** "Properties of elements are periodic function of their atomic masses."

Periodic table: "Properties of elements are periodic function of their atomic number."

20. Atomic number is a more fundamental property than atomic mass. Give reason. (DGK. GI)

Ans. Atomic number is the basic property, instead of atomic mass, because it is specific for every element while atomic mass may be same of two elements due to their isotopes.

21. How many blocks are present in periodic table? (DGK. GI)

Ans. In periodic table, there are four blocks s, p, d, f are present.

22. How many elements are found in the first period and what are their names? (DGK. GI)

Ans. In first period, only two elements are found hydrogen and helium.

23. Write down the number of group and periods in the long form of periodic table. (DGK. GII)

Ans. Long periodic table consist of 18 group and 7 periods.

3.2 Periodicity of Properties

☆ Tick the correct answer.

1. The electronegativity of carbon is: (LHR. GI, MLN. GII)

(A) 2.0 (B) 1.0 (C) 2.6 (D) 4.0

2. The distance between the nuclei of two carbon atoms is: (LHR. GI, MLN. GII)

(A) 154 Pm (B) 140 Pm (C) 110 Pm (D) 115 Pm

3. Which one of the following halogen has highest electronegativity? (LHR. GII)

(A) Iodine (B) Bromine (C) Chlorine (D) Fluorine

4. Electronegativity of oxygen is: (GRW. GII)

(A) 3.1 (B) 3.3 (C) 3.2 (D) 3.4

5. The electronegativity of nitrogen is: (FBD. GII, BWP. GI, FBD. GI)

(A) 2 (B) 3 (C) 4 (D) 5

6. **The amount of energy given out when an electron is added to an atom is called:** (MLN. GI, SWL. GI, SGD. GI)
(A) Lattice Energy (B) Ionization Energy
(C) Electronegativity (D) Electron Affinity

7. **Which one of the following halogens has the lowest electronegativity.** (SWL. GI, LHR. GI, GRW. GI, SWL. GII, SGD. GI, DGK. GII, BWP. GII)
(A) fluorine (B) chlorine (C) iodine (D) bromine

8. **The electronegativity of fluorine.** (SWL. GII, BWP. GII)
(A) 4.0 (B) 3.5 (C) 2.1 (D) 3.0

9. **The radius of Carbon atom is:** (SGD. GII, BWP. GI, GRW. GI)
(A) 154 pm (B) 115 pm (C) 77 pm (D) 38 pm

10. **Which element has the least value of electronegativity?** (GRW. GII)
(A) lithium (B) beryllium (C) boron (D) carbon

11. **The difference between electronegativity of hydrogen and chlorine is:** (FBD. GI)
(A) 1.0 (B) 1.6 (C) 1.8 (D) 2.0

12. **_____ has least value of shielding effect:** (MLN. GII)
(A) Lithium (B) Sodium (C) Potassium (D) Rubidium

13. **Atomic size of sodium atom is:** (SGD. GII)
(A) 160 Pm (B) 162Pm (C) 185Pm (D) 186Pm

14. **Which one of the following decreases in periods of periodic table:** (SGD. GII)
(A) atomic radius (B) ionization energy
(C) electron affinity (D) dative covalent bond

15. **Point out the incorrect statement about electron affinity:** (BWP. GI)
(A) it is measured in $KJmol^{-1}$ (B) it decreases in period
(C) it involves release of energy (D) it decrease in a group

16. **Ionization energy increases in period because:** (BWP. GI)
(A) number of shells increases (B) number of shell decreases
(C) number of electrons decreases
(D) force of attraction between valence shell electrons and nucleus increases

Answers

1.	2.6	**2.**	154 Pm	**3.**	Fluorine	**4.**	3.4
5.	3	**6.**	Electron Affinity	**7.**	iodine	**8.**	4.0
9.	77 pm	**10.**	lithium	**11.**	1.0	**12.**	Lithium
13.	186Pm	**14.**	atomic radius	**15.**	it decreases in period		

16. force of attraction between valence shell electrons and nucleus increases

☆ Give short answer to the following questions.

1. Why size of atom increases from top to bottom in periodic table? (LHR. GI, SWL. GII)

Ans. The size of atom increases down the group due to increase in number of shells of electron. Due to this, the effective unclear charge become decreased.

2. Why noble gases are not reactive? (LHR. GI, BWP. GI & II, MLN. GI, SWL. GII)

Ans. Nobel gases have 2 or 8 electrons in their valence shell. It means their valence shells are complete, and there is no space for more electrons. Due to this stability they neither gain or loss electron, nor they share electrons. That is why they are non reactive.

3. State about the trend of ionization energy in a period. (LHR. GII, FBD. GII)

Ans. In period, ionization energy increases from left to right. The reason is reduced size of atom due to more electrostatic force of nucleus on valence electrons.

4. Define atomic radius. (LHR. GII, MLN. GII, SWL. GII, RWP. GI)

Ans. The half of the distance between nuclei of two bonded atoms is called atomic radius. It is measured in peco meter (pm = 10^{-12}m)

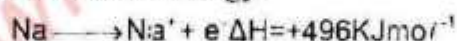
5. Define electron affinity with an example. (GRW. GI, SWL. GI, SGD. GI)

Ans. The amount of energy released when an electron is added to an isolated gaseous atom is called electron affinity.

$$F + e^- \longrightarrow F^- \quad \Delta H = -328 kJmol^{-1}$$

6. Define ionization energy with an example. (GRW. GI, SWL. GI, RWP. GII)

Ans. The amount of energy required to remove the outermost electron from an isolated gaseous atom is called ionization energy.

$$Na \longrightarrow Na^+ + e^- \Delta H = +496 KJmol^{-1}$$

7. Write trend of electronegativity in a group. (GRW. GII)

Ans. Electronegativity decreases down the group. Because atomic size increases resulting in less force of attraction of nucleus toward valence electrons.

8. What is meant by periodic function? (GRW. GII, BWP. GII)

Ans. Those properties which are repeated after specific interval of time, known as periodic function. e.g chemical properties, electronic configuration.

9. What is shielding effect? (GRW. GII, DGK. GI & II, GRW. GII, RWP. GI, BWP. GI)

Ans. The decrease in attractive force exerted by nucleus on valence electrons due to presence of electrons lying between nucleus and valence shell electron is called shielding effect.

10. What is the trend of electron egativity in period? (FBD. GI)

Ans. Along the period, electronegativity increased. In period, the number of proton in nucleus increased, resulting in more attraction of nucleus toward valence electron, so the size of atom become small.

11. Why shielding effect of electrons makes cation formation easy? (FBD. GI)

Ans. When the shielding effect of electron increases, it decreases the attractive force between valence electron and nucleus, which makes the removal of electron easy.

12. How is periodicity of properties dependent upon the number of protons in an atom? (FBD. GI)

Ans. The properties of atom depend upon its size. And the size of atom depend the Z-effect or nuclear charge, which is due to numbers of protons. Greater the number of protons, greater will be nuclear charge, and results in small size of atom.

13. What is the trend of ionization energy in a group and period? (FBD. GII)

Ans. Along the period, from left to right ionization energy increases. While along group, from top to bottom ionization energy decreased.

14. Why the size of Atoms decrease in a Period? (MLN. GI, RWP. GII, GRW. GI)

Ans. The reason is along the period, number of protons in nucleus due to increase in atomic number increases, resulting in more nuclear charge toward revolving electrons. But the number of shells do not increases. So there is more attraction of nucleus toward valence shell's electron that's why atomic size reduced.

15. Why the size of atom does not decrease regularly in a period? (SWL. GI)

Ans. Due to weak shielding effect, the size of atom do not regularly reduced.
Along the period firstly the atomic size decrease than increases.

16. Why the second ionization energy of an element is higher than first one? (LHR. GI)

Ans. The second ionization is greater because when an electron is removed, the number of electrons become decreased while number of protons will remain same, so the size of atom decreases, which increased the ionization energy.

17. Define Shielding Effect. Describe its trend along the period. (LHR. GII)

Ans. The decrease in attractive force of nucleus toward valence electrons due to presence of electrons in inner shells. This is known as shielding effect.
Trend: Along the period it is decreased.

18. Give the trend of electron affinity in periods. (FBD. GII)

Ans. The value of electron affinity increase along the period. The reason is that, the atomic size get reduced along period, resulting in attraction of nucleus toward valence electron. It means, more the attraction will be, the greater is energy released, by adding electrons.

19. Define Electronegativity. (MLN. GI, DGK. GII)

Ans. Ability of atom to attract shared pair of electron toward itself is known as electronegativity.

STRUCTURE OF MOLECULES

Major Concepts:

4.1 Why do atoms react? 4.2 Chemical bonds
4.3 Types of bonds
4.4 Intermolecular forces
4.5 Nature of bonding and properties

Time allocation	
Teaching periods	16
Assessment periods	04
Weightage	8%

Students Learning Outcomes:

Students will be able to:

- Find the number of valence electrons in an atom using the Periodic Table.
- Describe the importance of noble gas electronic configurations.
- State the octet and duplet rule.
- Explain how elements attain stability.
- Describe the ways in which bonds may be formed.
- State the importance of electronic configurations in formation of ion.
- Describe formation of cations from an atom of a metallic element.
- Describe formation of anion from a non-metallic element.
- Describe characteristic of ionic bond.
- Recognize a compound as having ionic bonds.
- Identify characteristics of ionic compounds.
- Describe formation of covalent bond between two non-metallic elements.
- Describe with examples single, double and triple covalent bonds.
- Draw electron cross and dot structure of simple covalent molecules containing single, double and triple covalent bonds.

4.1 WHY DO ATOMS FORM CHEMICAL BONDS?

4.2 CHEMICAL BONDS

Q1. (a) What is meant by chemical bond? Why atoms form chemical bond.
(b) Define the following. (i) Duplet rule. (ii) Octet rule.

Ans:(a) Chemical bond: The attractive force which binds the particles [atoms, ions or molecules] together is called a chemical bond:

Why atoms form chemical bond? Atoms combine with each other to form bond because they want to decrease their energy and to get stability. It is natural tendency that

every thing in this universe wants to get stability by decreasing its energy. Atoms decrease their energy by completing their valence shells.

All the atoms that are present in the periodic table are unstable because of their incomplete valence shells except group VIII-A elements. These elements [group VIII - A] have complete valance shells and have little tendency to combine with other elements hence are called noble or inert gases.

Atoms of all other elements [except noble gases] try to get noble gas configuration. This tendency is an important basis of the chemical bonding.

(b) (i) Duplet rule:

The ability of an atom to get two electrons in its valence shell is called duplet rule.

(ii) Octet rule: The ability of an atom to get eight electrons in its valence shell is called octet rule.

4.3 TYPES OF CHEMICAL BOND

Q2. (a) What is meant by valence electrons. Write down the names of types of chemical bond.

(b) What is meant by ionic or electrovalent bond? Explain the formation of ionic bond with example.

Ans:(a) Valence electrons:

The valence electrons, which are involved in chemical bonding, are termed as bonding electrons. They usually reside in the incomplete or partially filled outer most shell of an atom.

Types of chemical bond: A chemical bond has following types.

1: Ionic Bond
2: Covalent Bond
3: Dative covalent or coordinate covalent Bond
4: Metallic Bond

(b) Ionic or Electrovalent bond: A chemical bond, which is formed by the electrostatic forces of attraction is called ionic or electrovalent bond.

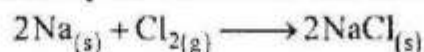
OR

The chemical bond which is formed by the complete transfer of one or more electrons from one atom to another atom is called ionic bond.

Examples of ionic bond: Ionic bond has following examples.

(i) Formation of Sodium chloride:

Sodium chloride is formed by the reaction of sodium with chlorine.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

Sodium atom loses one electron to get noble gas electronic configuration and becomes Na^{+} i.e

$$\underset{(1s^2, 2s^2 2p^6, 3s^1)}{\overset{(2,8,1)}{Na}} \longrightarrow \underset{(1s^2, 2s^2, 2p^6)}{\overset{(2,8)}{Na^{+1}}} + 1e^-$$

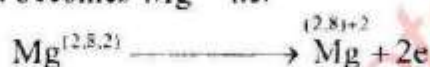
Similarly chlorine atom gains one electron to get noble gas electronic configuration and becomes Cl^- i.e.

$$\underset{(1s^2,2s^2,2p^6,3s^2,3p^5)}{\overset{(2,8,7)}{Cl}} + e^- \longrightarrow \underset{(1s^2,2s^2,2p^6,3s^2,3p^6)(Ar)}{\overset{(2,8,8)}{Cl^-}}$$

These oppositely charged ions $[Na^+, Cl^-]$ formed the ionic bond between sodium and chlorine.

$$Na^+ + Cl^- \longrightarrow NaCl$$

(ii) Formation of Magnesium oxide: Magnesium oxide is formed by the reaction of Magnesium with oxygen. Magnesium atom loses two electrons to get noble gas electronic configuration and becomes Mg^{2+} i.e.

$$Mg^{(2,8,2)} \longrightarrow \overset{(2,8)+2}{Mg} + 2e^-$$

Similarly oxygen atom gains two electrons to get noble gas electronic configuration and becomes O^{-2} i.e.

$$O^{(2,6)} + 2e^- \longrightarrow O^{-2(2,8)}$$

These oppositely charged ions $[Mg^{2+}, O^{2-}]$ form the ionic bond between magnesium and oxygen.

$$Mg^{+2} + O^{-2} \longrightarrow MgO$$

Test your self 4.1:

i. Why does sodium form a chemical bond with chlorine?

Ans. Sodium forms chemical bond with chlorine because sodium has one electron in its last shell and chlorine has seven electrons in its last shell. Sodium loses one electron and becomes Na^{+1} and chlorine gains one electron and becomes Cl^{-1}. These oppositely charged ions causes the ionic bond between sodium and chlorine.

ii. Why does sodium lose an electron and attains +1 charge?

Ans. Sodium loses one electron and becomes Na^{+1} inorder to get noble gas electronic configuration.

iii. How do atoms follow octet rule?

Ans. The ability of an atom to get eight electrons in the last shell is called octet rule. The atoms having 1, 2 or 3 electrons lose electrons to get noble gas electronic configuration while the atoms having 4-7 electrons gains electrons to complete their octet.

iv. Which electrons are involved in chemical bonding?

Ans. Valence electrons are involved in chemical bonding.

v. Why does group I elements prefer to combine with group 17 elements.

Ans. Group one elements having only one electron in their valence shell. They lose this one electron and become uninegative ions. The group 17 elements are required only one electron so they gain this electron to become uninegative ion. These oppositely charged ions cause the ionic bond between group I and group 17.

vi. Why chlorine can accept only 1 electron?

Ans. Chlorine belongs to group 17 and has seven electrons in its last shell. It accept one electron to complete its octet.

Q3. (a) What is meant by covalent bond? Explain the formation of covalent bond with examples.

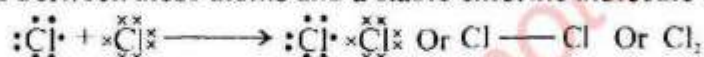
(b) Describe the types of covalent bond with examples.

Ans:(a) A chemical bond formed by the mutual sharing of electrons is called covalent bond.

Formation of covalent bond: A covalent bond is usually formed between two non metal atoms. Both atoms contribute equal number of electrons in the formation of covalent bond. By sharing electrons both atoms attain the noble gas electronic configuration. The two atoms are linked together due to the attraction between the shared electrons and the positive nuclei of the atoms.

Example:

Formation of chlorine molecule (Cl_2): When two atoms of chlorine combine, a pair of electrons is shared between these atoms and a stable chlorine molecule is formed:

$$:\ddot{\underset{..}{Cl}}\cdot + \times\overset{\times\times}{\underset{\times\times}{Cl}}\overset{\times}{\times} \longrightarrow :\ddot{\underset{..}{Cl}}\cdot\times\overset{\times\times}{\underset{\times\times}{Cl}}\overset{\times}{\times} \text{ Or } Cl-Cl \text{ Or } Cl_2$$

(b) Types of covalent bond: A covalent bond has following types:

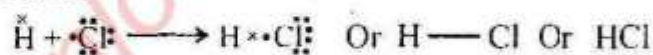
(i) Single covalent bond (ii) Double covalent bond
(iii) Triple covalent bond

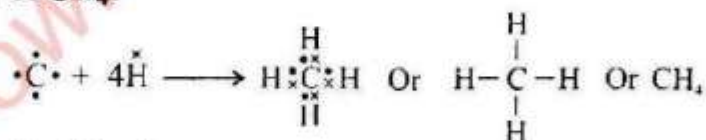
(i) Single covalent bond:

A covalent bond which is formed by the mutual sharing of one electron pair is called single covalent bond. It is represented by single short straight line (—).

Examples:

(i) Formation of HCl:

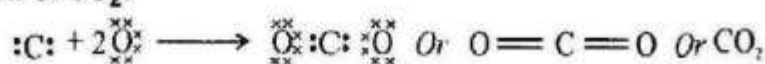
$$\overset{\times}{H} + \cdot\ddot{\underset{..}{Cl}}: \longrightarrow H\times\cdot\ddot{\underset{..}{Cl}}: \text{ Or } H-Cl \text{ Or } HCl$$

(ii) Formation of CH_4:

$$\cdot\dot{\underset{\cdot}{C}}\cdot + 4\overset{\times}{H} \longrightarrow H\overset{\times}{\underset{\times}{:}}\overset{H}{\underset{H}{\overset{\times\times}{\underset{\times\times}{C}}}}\overset{\times}{\times}H \text{ Or } H-\overset{H}{\underset{H}{\overset{|}{\underset{|}{C}}}}-H \text{ Or } CH_4$$

(2) Double covalent bond:

A covalent bond which is formed by the mutual sharing of two pairs of electrons is called double covalent bond. It is represented by two short straight lines. (═)

(i) $$\ddot{\underset{..}{O}}: + \overset{\times}{\underset{\times}{:}}\overset{\times\times}{\underset{\times\times}{O}} \longrightarrow \ddot{\underset{..}{O}}:\overset{\times}{\underset{\times}{:}}\overset{\times\times}{\underset{\times\times}{O}} \text{ Or } O=O \text{ Or } O_2$$

(ii) Formation of CO_2:

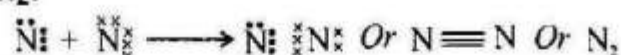
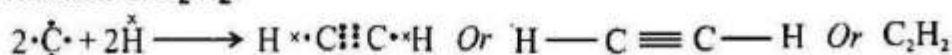
$$:C: + 2\overset{\times\times}{\underset{\times\times}{O}}\overset{\times}{\times} \longrightarrow \overset{\times\times}{\underset{\times\times}{O}}\overset{\times}{\times}:C:\overset{\times}{\times}\overset{\times\times}{\underset{\times\times}{O}} \text{ Or } O=C=O \text{ Or } CO_2$$

Triple covalent bond:

A chemical bond which is formed by the mutual sharing of three pairs of electrons is called triple covalent bond. It is represented by three short straight lines. (≡)

Examples:

(i) Formation of N_2:

$$\ddot{N}: + \ddot{N}\dot{\dot{x}} \longrightarrow \ddot{N}:\,:N: \;Or\; N\equiv N \;Or\; N_2$$

(ii) Formation of C_2H_2:

$$2\cdot\dot{C}\cdot + 2\overset{x}{H} \longrightarrow H\times\cdot C::C\cdot\times H \;Or\; H-C\equiv C-H \;Or\; C_2H_2$$

Do you know?

The electronic configuration of the valence shells of atoms is shown in small 'dots' or 'crosses' around the symbol of the element. Each dot or cross represents an electron. This is a standard method of Lewis to describe the electronic configuration of valence shell of an atom. It is called Lewis Structure Diagram.

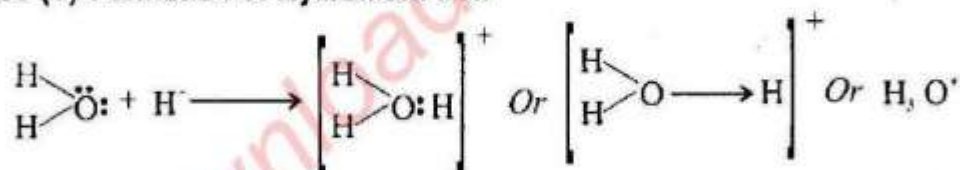
Q4. What is meant by co-ordinate covalent bond? Explain with examples?

Ans: Co-ordinate covalent bond: A covalent bond in which the shared pair of electrons is donated the one atom only is called a co-ordinate covalent bond.

Formation of co-ordinate covalent bond: A co-ordinate covalent bond is formed, when a molecule has an electron pair which can be donated to another molecule. The molecule which donates the electron pair is called donor and which accepts the electron pair is called an acceptor.

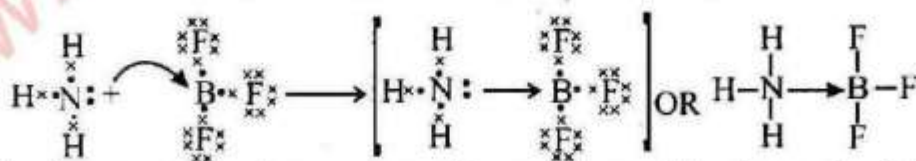
Representation: The co-ordinate covalent bond is represented by an arrow pointing toward the acceptor.

Examples (1) Formation of Hydronium Ion:

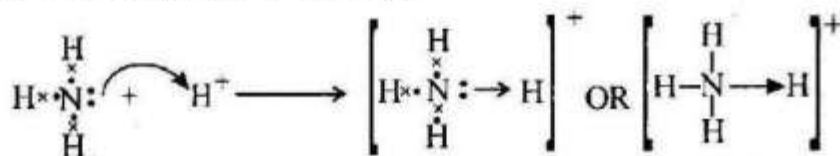
$$H_2\ddot{O}: + H^+ \longrightarrow \left[H_2O:H\right]^+ \;Or\; \left[H_2O\longrightarrow H\right]^+ \;Or\; H_3O^+$$

In above reaction oxygen atom of water molecule donates a lone pair of electrons to form a co-ordinate covalent bond with hydrogen ion.

(2) Reaction between NH_3 and BF_3: The reaction between ammonia (NH_3) and Boron (Trifluoride (BF_3) is an example of co-ordinate covalent bond.

$$H_3N: + BF_3 \longrightarrow H_3N \longrightarrow BF_3$$

$$H_3N: + BF_3 \longrightarrow \left[H_3N:\longrightarrow BF_3\right] \;OR\; H_3N\longrightarrow BF_3$$

In the above reaction nitrogen atom of ammonia provide a lone pair of electrons to the boron atom of borontrifluoride to form a co-ordinate covalent bond.

(3) Formation of ammonium ion NH_4^+:

$$H_3N: + H^+ \longrightarrow \left[H_3N:\rightarrow H\right]^+ \;OR\; \left[H_3N\longrightarrow H\right]^+$$

In the above reaction the non bonded electron pair of ammonia is donated to H^+ and a coordinate covalent bond is formed.

Q5. (a) Define the following:

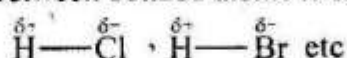
(i) Polar covalent bond ***(ii) Non polar covalent***

(iii) Polar molecules ***(iv) Non polar molecules***

(b) What is meant by metallic bond? Explain the formation of metallic bond.

Ans:(a) (i) Polar covalent bond:

Covalent bond which is formed between two dissimilar atoms having a reasonable difference of electronegativity between bonded atoms is called polar covalent bond e.g.

$$\overset{\delta+}{H} - \overset{\delta-}{Cl} \cdot \overset{\delta+}{H} - \overset{\delta-}{Br} \text{ etc}$$

(ii) Non polar covalent bond: Covalent bond which is formed between similar atoms having same electro negativity between bonded atoms is called non polar covalent bond.

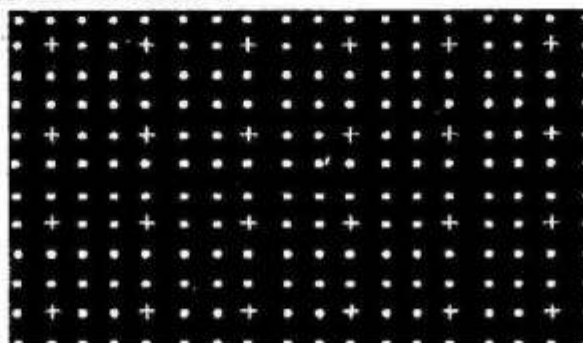
$$\overset{\delta+}{H} - \overset{\delta+}{H}$$

(iii) Polar molecules: The molecules or compounds which have polar covalent bonds are called polar molecules. e.g. $\overset{\delta+}{H} - \overset{\delta-}{Cl} \cdot \overset{\delta+}{H} - \overset{\delta-}{Br}$

(iv) Non polar molecules: The molecules which have non polar covalent bonds are called non-polar molecules e.g. H_2 , Cl_2 , etc.

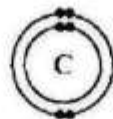
(b) Metallic bond: When positively charged metal ions are held together by freely moving electrons, the bond which is formed, called metallic bond.

When positively charged metal ions are held together by freely moving electrons, the bond formed is called metallic bond.

Formation of metallic bond: Metals have low ionization energy values. The metal atoms can easily lose their valence electrons. The nuclei of metal atoms can not hold the outer electrons firmly and they can move freely in vacant spaces present between atoms. No electrons remains attached with any particular atom. All the electrons are attached with all the atoms. When all the atoms will attract all the electrons collectively they will bound together, and a metallic bond is formed.

A schematic diagram of Copper wire showing its positive nuclei (+) embedded in sea of free electrons (o) making 'Metallic Bonding'

Test your self 4.2:

i. Give the electronic configuration of carbon atom.

Ans. Electronic configuration of carbon:

$1s^2, 2s^2, 2p^6$

ii. What type of elements have tendency of sharing of electrons?

Ans. Non-metals have tendency of sharing of electrons.

iii. If repulsive forces dominate to attractive forces will a covalent bond form?

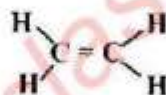
Ans. If repulsive forces do not dominate to attractive forces a covalent bond is not formed.

iv. Considering the electronic configuration of nitrogen atom, how many electrons are involved in bond formation and what type of covalent bond is formed.

Ans. Nitrogen atom uses three electrons to form triple covalent bond.

v. Point out the type of covalent bonds in the following molecules.
CH_4, C_2H_4, H_2, N_2 and O_2

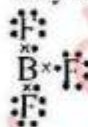
Ans. CH_4 = Single covalent bond:
C_2H_4:

H–C(–H)=C(–H)–H

Carbon to carbon double and carbon to hydrogen single covalent bonds.
H_2: Single covalent bond
N_2: Triple covalent bond
O_2: Dobule covalent bond

vi. What is a lone pair? How many lone pair of electrons are present on nitrogen in ammonia?

Ans. Lone pair: The pair of electrons which does not contribute in bonding is called lone pair of electrons. Nitrogen has only one lone pair of electrons in ammonia.

vii. Why is the BF_3 electron deficient?

Ans. :F: B ×· F: :F:

Boron has three electrons in its last shell. Three fluorine atoms shares one electron each to complete their octets. After this sharing boron has six electrons in its last shell. So it is deficient in electrons.

viii. What types of electron pairs make a molecule good donor?

Ans. Lone pair of electrons make a molecule good donor.

ix. What is difference between bonded and lone pair of electron and how many bonded pair of electrons are present in NH_3 molecule?

Ans. Bonded pair of electrons: The pair of electrons which is shared between the bonded atoms is called bonded pair of electrons.

Lone pair of electrons: The pair of electrons which does not contribute in bonding is called lone pair of electrons.

NH_3: Ammonia has there bonded pair of elecrons.

x. What do you mean by delta sign and why it develops?

Ans. The delta (δ) sign indicates partial positive or partial negative charge that is developed due to unequal sharing of shared pair of electrons.

xi. Why does oxygen molecule not form a polar covalent bond?

Ans. Oxygen forms non-polar covalent bond because shared pair of electrons is equally shared between two oxygen atoms having same electronegativity.

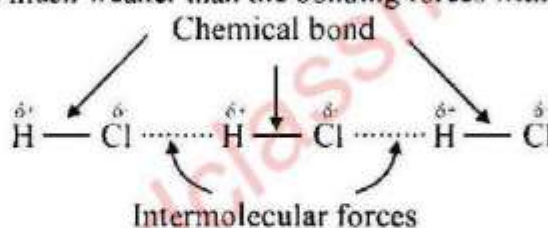
xii. Why has water polar covalent bonds?

Ans. Water forms polar covalent bonds due to the unequal attraction to the shared pair of electrons between oxygen and hydrogen atoms that having different electronegativity values.

4.4 INTERMOLECULAR FORCES

Q6. What is meant by intermolecular forces? Describe the various types of intermolecular forces with examples.

Ans: Intermolecular forces: The forces of attraction present between the molecules of a compound are called intermolecular forces.

These forces are much weaker than the bonding forces with in these molecules.

It requires about 17 KJ energy to break these intermolecular forces between one mole of liquid hydrogen chloride molecule to converted it into gas whereas about 430 KJ energy is required to break the chemical bond between hydrogen and chlorine atoms in 1 mole of hydrogen chloride.

Types of intermolecular forces: The intermolecular forces are collectively called vander wall's forces. Two types are discussed below.

(1) Dipole-dipole forces: The attractive forces between positive end of one polar molecule and negative end of other polar molecule are called dispole-dispole forces.

OR

The forces of attraction which originate due to the difference in electronegativities of bonded atoms in polar molecules are called dipole-dipole forces.

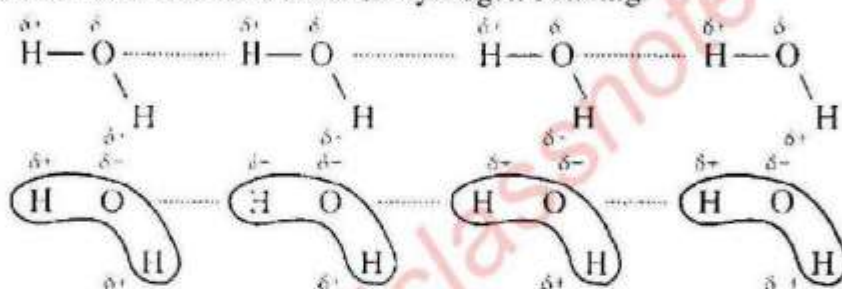
Explanation: Dipole-dipole forces are present between the molecules of polar compounds. In polar molecules one end of the molecule has partial positive (δ+) charge while other end of the molecule has partial negative (δ-) charge due to the difference in electro negativities of the bonded atoms. When polar molecules come closer to each other, they arrange themselves in such a way that positive end of one molecule attracts the negative end of other molecule. Thus a force of attraction is created, which is called dipole-dipole interaction e.g. HCl is a polar molecular due to the difference in electro-negativities of the bonded atoms. The polar molecules of HCl attract each other with a force, which is called dipole-dipole force.

$$\overset{\delta+}{H} - \overset{\delta-}{Cl} \cdots\cdots \overset{\delta+}{H} - \overset{\delta-}{Cl} \cdots\cdots \overset{\delta+}{H} - \overset{\delta-}{Cl}$$

Polar compounds have high melting and boiling points due to the presence of dipole - dipole forces.

(2) Hydrogen bonding:

The electrostatic force of attraction between covalently bonded polarized hydrogen atom and any other more electronegative atom (F, O, N) is called hydrogen bonding, it is represented by dotted line. It is present in compounds in which the hydrogen atom is bonded to any one of the very strong electronegative atom (F, O, N).

Consider the water molecule to understand the hydrogen bonding, oxygen atom is more electronegative than hydrogen atom so water is a polar molecule. There will be a force of attraction between polar water molecules. These electrostatic force of attraction between electronegative oxygen atom of one molecule and partial positive charge hydrogen atom of other molecule is called hydrogen bonding.

Test your self 4.3:

i. What type of elements form metallic bonds?

Ans. Metals form metallic bonds.

ii. Why is the hold of nucleus over the outermost electrons in metals weak?

Ans. In case of metals, the hold of nucleus over the outermost electrons is weak because of large sized atoms and greater number of shells in between nucleus and valence electrons.

iii. Why the electrons move freely in metals?

Ans. Free electrons of all the metal atoms move freely in the spaces between atoms of metals. None of these electrons are attached to any particular atom, either they belong to a common pool or belong to all the atoms of metal hence they move freely.

iv. Which types of electrons are responsible for holding the atoms together in metals.

Ans. Mobile electrons are responsible for holdings the atoms together in metals.

v. Why a dipole develops in a molecule?

Ans. The dipole develops in molecule due to the unequal attraction to the shared pair of electrons between bonded atoms e.g.

$$\overset{\delta+}{H}-\overset{\delta-}{Cl}$$

vi. What do you mean by induced dipole?

Ans. The pole which is produced in a non-polar molecule due to a polar molecule is called induced dipole.

vii. Why are dipole forces of attraction not found in halogen molecules?

Ans. Because they are nonpolar.

viii. What types of attractive forces exist between HCl molecules?

Ans. Dipole-dipole interaction.

ix. Define intermolecular forces; show these forces among HCl molecule.

Ans. The forces of attraction present between the two molecules are called intermolecular forces.

$$\overset{\delta+}{H}-\overset{\delta-}{Cl}\cdots\overset{\delta+}{H}-\overset{\delta-}{Cl}\cdots\overset{\delta+}{H}-\overset{\delta-}{Cl}$$

4.5 Nature of Bonding and Properties.

Q7. Differentiate the ionic and covalent compounds.

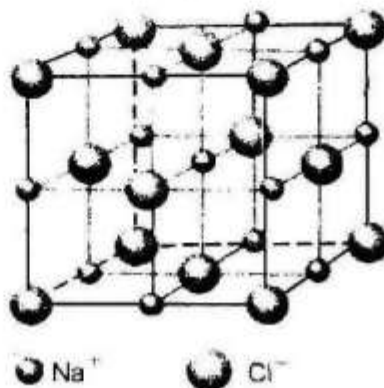
Ans:

Ionic Compounds	Covalent Compounds
1- The compounds which contain ionic bonds are called ionic compound e.g. NaCl, MgO etc.	1- The compounds which contain covalent bonds are called covalent compounds e.g. H_2O, HCl etc
2- They consist of positive and negative ions.	2- Most of the covalent compounds consist of molecules.
3- They exist only in the solid state at room temperature.	3- They exist as solid, liquid or gas at room temperature.
4- They have high melting and boiling points.	4- They have low melting and boiling points.
5- They are soluble in polar solvents e.g. water.	5- They are mostly soluble in non-polar solvents e.g. benzene.
6- They are non-volatile.	6- They are volatile.
7- They are good conductor of heat and electricity in aqueous form or in the fused state.	7- Pure covalent compounds do not conduct electricity.
8- They are hard.	8- They are usually soft.

Q8. What are ionic compounds? Draw the structure of NaCl. Write properties of NaCl also.

Ans: Ionic compound: The compounds which contain ionic bonds are called ionic compounds e.g. NaCl.

Ionic compounds are made up of positively and negatively charged ions. They consist of ions not molecules.

Structure of NaCl: Sodium and chloride ions are held together in a solid crystal of sodium chloride are held together by strong electrostatic forces of attraction. The orderly arrangement of Na^+ and Cl^- ions in a solid crystal of sodium chloride is shown.

Regular arrangement of Na^+ and Cl^- ions in solid crystal of NaCl

Properties of ionic compounds: Ionic compounds having following properties.

(1) Ionic compounds are mostly crystalline solids.
(2) Ionic compounds are good conductors of electricity in the molten or in aqueous solution but are non-conductors in solid state.
(3) Ionic compounds have high melting and boiling points e.g melting point of NaCl is 800°C.
(4) Ionic compounds are usually soluble in water and in polar solvents [water has high electric constant that weaken the attraction between ions].
(5) Ionic compounds have low density.

Q9. What are covalent compounds? Write down their properties.

Ans: The covalent compound are made up of molecule that are formed by mutual sharing of electrons between their atoms.

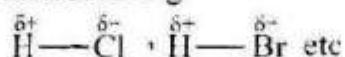
Covalent compounds: The compounds having covalent bonds are called covalent compounds. They are made up of molecules. e.g. CH_4, H_2O, HCl.

Properties of covalent compounds: Covalent compounds have following properties.

(1) Covalent compounds have low melting and boiling points.
(2) Pure covalent compounds are bad conductors of electricity but some polar covalent compounds are good conductors of electricity in aqueous form.
(3) Pure covalent compounds are insoluble in water but are soluble in non-aqueous solvents. Like benzene.
(4) Bigger covalent molecules are stable and hard and have high melting and boiling points.
(5) Covalent compounds usually have low density.

Q10. (a) What is meant by polar and non-polar compounds. Give examples.
(b) Differentiate the polar covalent and non-polar covalent compounds.

Ans:(a) Polar covalent compounds: The molecules or compounds which have polar covalent bonds are called polar molecules e-g

$$\overset{\delta+}{H}—\overset{\delta-}{Cl} \;,\; \overset{\delta+}{H}—\overset{\delta-}{Br} \text{ etc}$$

Non-polar covalent molecules: The molecules which have non-polar covalent bonds are called non-polar molecules e-g. H_2, Cl_2 etc.

(b) Differences between polar and non-polar covalent compound:

Polar covalent	Non-polar covalent
1- Polar covalent compounds are soluble in water.	1- Non-polar covalent compounds are insoluble in water.
2- Polar covalent compounds. usually conduct electricity in aqueous form.	2- Non-polar covalent compounds do not conduct electricity.

Q11. (a) Write down the properties of metals.
(b) Why Ice floats on the surface of water.
(c) Define coordinate covalent compounds.

Ans: (a) Properties of metlas: Major properties shown by the metals are as following.

(i) They show metallic luster.
(ii) They are usually malleable and ductile. Malleability is the property by virtue of which a metal can be rolled into sheets, while ductility is the property by virtue of which a metal can be drawn into wires.
(iii) They have usually high melting and boiling points.
(iv) Being greater in size they have low ionization energies and form cations (M^+) very easily.
(v) They are good conductors of heat and electricity in solid and liquid state due to mobile electrons.

(b) Ice floats on the surface of water:

The ice floats at the surface of water due to the presence of hydrogen bonding. The density of ice at 0°C is less than that of liquid water. In liquid state, the water molecules move randomly. However when water freezes, The molecules arrange themselves in ordered form that gives them open structure. When the ice is formed at 0°C it expands still further, so ice is less dense than water and floats on it.

(c) Define Coordinate covalent compounds.

When a fixed number of molecules or ions are bonded to a metal atom through co-ordinate covalent bonds, the compounds are called co-ordinate covalent compounds.

Test your self 4.4:

i. Why the ionic compounds have high melting and boiling points?

Ans. Ionic compounds have high melting and boiling points due to strong electrostatic forces of attraction between oppositely charged ions.

ii. What do you mean by malleability?

Ans. The property of metals by virtue of which metals can be drawn into sheets.

iii. Why are ionic compounds easily soluble in water?

Ans. Ionic compounds are soluble in water because water has high dielectric constant which weakens the attractions between ions and cause them to dissolve.

iv. What type of bond exists in sodium chloride?

Ans. Ionic or electrovalent bond, exists in sodium chloride.

v. Why the covalent compounds of bigger size molecules have high melting points?

Ans. Covalent compounds of bigger size have high melting and boiling points due to the very large number of bonds, which make them stable.

vi. (a) What is the electronegativity difference between the following pair of elements (atoms). Predict the nature of the bond between them?

(a) H and Cl (b) H and Na (c) Na and I (d) K and Cl

(b) Comparing the electronegativity differences, arrange these compounds in increasing ionic strength.

Ans. (a) H and Cl

$3.2 - 2.2 = 1.0$

covalent

H and Na

$2.2 - 0.9 = 1.3$ Ionic

(c) Na and I

$2.7 - 0.9 = 1.8$ Ionic

(d) K and Cl

$3.2 - 0.8 = 2.4$ Ionic

(b) KCl < NaI < NaH < HCl

Synthetic Adhesives:

Although natural adhesives are less expensive to produce, but most important adhesives used now a days are synthetic. Adhesives based on synthetic resins and rubbers excel in versatility and performance, Synthetic adhesives can be produced in a sufficient supply with uniform properties and they can be modified in many ways. They polymers or resins used in synthetic adhesives fall into two general categories - thermoplastics and thermosetting. One form of polymer used industrially is epoxy adhesive.

AIR CRAFTS, CARS, TRUCKS AND BOATS ARE PARTIALLY HELD TOGETHER WITH EPOXY ADHESIVES:

Epoxy is polymer that is formed from two different chemicals. These are referred to as resin and the hardener. Epoxy adhesives are called structural adhesives. These high-performance adhesives are used in the construction of aircraft, automobiles, bicycles, boats golf clubs, where high strength bonds are required. Epoxy adhesives can be developed to suit almost any application. They can be made flexible or rigid, transparent or opaque even colored as well as fast or slow setting. Epoxy adhesives are good heat and chemical resistant. Because of these properties, they are given the name of engineering adhesives.

Key Points

- Atoms of different elements react to attain noble gas configuration, which is stable one.
- Chemical bonds may be formed by complete transfer of electrons (ionic); mutual sharing (covalent) or by donation from an atom (coordinate or dative covalent).
- Metals have the tendency to lose electrons easily forming cations.
- Non-metals have tendency to gain electron and form anions.
- In ionic bonding strong electrostatic force hold ions together.
- Ionic compounds are solids with high melting and boiling points.
- Covalent bonds among non-metals are weaker than ionic bonds.
- Ionic bonds are non-directional, but covalent bonds are formed in a particular direction.
- Covalent bonds formed between similar atoms are non-polar while between different atoms are polar.
- In covalent bonding single, double or triple covalent bond is formed by sharing of one, two or three electron pairs by the bonded atoms.
- Coordinate covalent bond is formed between electron pair donors and electron pair acceptors.
- Metallic bond is formed between metals due to free electrons.
- In addition to chemical bonds, intermolecular forces of attraction exist between polar molecules.
- Hydrogen bonding exists between the hydrogen atom of one molecule and highly electronegative atom of other molecule.

- Hydrogen bonds affect the physical properties of the compounds.
- Properties of the compounds depend upon the nature of bonding present in the compound.
- Ionic compounds are crystalline solid with high melting and boiling points.
- Covalent compounds exist in molecular form in three physical states.
- Polar and non-polar covalent compounds have different properties.
- Metals have shining surface. They are good conductor of electricity and are malleabie and ductile.

Exercise (Solved)

Multiple Choice Questions

Put a (✓) on the correct answer.

1. Atoms react with each other because:
(a) they are attracted to each other (b) they are short of electrons
(c) they want to attain stability (d) they want to disperse

2. An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:
(a) gaining one electron (b) losing all electrons
(c) gaining two electrons (d) losing two electrons

3. Considering the electronic configuration of atoms which atom with the given atomic number will be the most stable one?
(a) 6 (b) 8 (c) 10 (d) 12

4. Octet rule is:
(a) description of eight electrons (b) picture of electronic configuration
(c) pattern of electronic configuration (d) attaining of eight electrons

5. Transfer of electrons between atoms results in:
(a) metallic bonding (b) ionic bonding
(c) covalent bonding (d) coordinate covalent bonding

6. When an electronegative element combines with an electropositive element the type of bonding is:
(a) covalent (b) ionic (c) polar covalent (d) coordinate covalent

7. A bond formed between two non-metals is expected to be:
(a) covalent (b) ionic (c) coordinate covalent (d) metallic

8. A bond pair in covalent molecules usually has:
(a) one electron (b) two electrons (c) three electrons (d) four electrons

9. Which of the following compounds is not directional in its bonding?
(a) CH_4 (b) KBr (c) CO_2 (d) H_2O

10. Ice floats on water because:
(a) ice is denser than water (b) ice is crystalline in nature
(c) water is denser than ice (d) water molecules move randomly

11. Covalent bond involves the:
(a) donation of electrons (b) acceptance of electrons
(c) sharing of electrons (d) repulsion of electrons

12. How many covalent bonds does C_2H_2 molecule have?
(a) two (b) three (c) four (d) five

13. How many electrons does a triple covalent bond involve?
(a) eight (b) six (c) four (d) only three

14. Which pair of the molecules has same type of covalent bonds?
(a) O_2 and HCl (b) O_2 and N_2 (c) O_2 and C_2H_4 (d) O_2 and C_2H_2

15. Identify the compound which is not soluble in water.
(a) C_6H_6 (b) NaCl (c) KBr (d) $MgCl_2$

16. Which one of the following is an electron deficient molecule?
(a) NH_3 (b) BF_3 (c) N_2 (d) O_2

17. Identify which pair has polar covalent bonds.
(a) O_2 and Cl_2 (b) H_2O and N_2 (c) H_2O and C_2H_2 (d) H_2O and HCl

18. Which one of the following is the weakest force among the atoms?
(a) ionic force (b) metallic force
(c) intermolecular force (d) covalent force

Answers:

1.	they want to attain stability	2.	gaining two electrons	3.	10
4.	attaining of eight electrons	5.	ionic bonding	6.	ionic
7.	covalent	8.	two electrons	9.	KBr
10.	water is denser than ice	11.	sharing of electrons	12.	five
13.	six	14.	O_2 and C_2H_4	15.	C_6H_6
16.	BF_3	17.	H_2O and HCl	18.	intermolecular force

Short Answer Questions.

Q1. Why do atoms react?

Ans: Atoms react with each other to complete their last shells to get stability.

Q2. Why is the bond between an electropositive and an electronegative atom ionic in nature?

Ans: The bond between an electropositive and an electronegative atom is ionic in nature because electropositive atom loses the electron to form a positive ions and electronegative atom gain electrons to give negative ions. These oppositely charged ions form an ionic bond.

Q3. Ionic compounds are solids. Justify.

Ans: Ionic compounds are solids because ionic compound, have very strong ionic bonds, combine ions to give solids.

Q4. More electronegative elements can form bonds between themselves. Justify.

Ans: Electronegative elements have ability to get electrons because magnetic ions combine with positive ions to form ionic bond.

Q5. Metals are good conductor of electricity. Why?

Ans: Metals are good conductors of electricity due to the presence of free electrons.

Q6. Ionic compounds conduct electricity in solution or molten form. Why?

Ans: Ionic salt conducts electricity in solution or molten form because they have free ions in solution or molten forms, and conduct electricity.

Q7. What type of covalent bond is formed in nitrogen molecule?

Ans: Triple covalent bond is formed in nitrogen molecule.

Q8. Differentiate between lone pair and bond pair of electrons.

Ans:

Lone Pair	Bond Pair
The pair electron which does not contribute in bonding and available on atom is called lone pair e.g. in NH_3, nitrogen atom has lone pair of electron. H:N̈:H H Lone Pair H—N̈—H | H	The shared pair of electrons between two bonded atoms is called bond pair of electrons e.g. bond pair H : H

Q9. Describe at least two necessary conditions for the formation of a covalent bond.

Ans: For answer See QNo.3(a).

Q10. Why HCl has dipole-dipole forces of attraction?

Ans: HCl has dipole-dipole force due to the unequal sharing of electrons between two different types of atoms make one end of molecule slightly positive and other end is slightly negative.

When partial positive and partial negative charges exist at different positions in a molecule, the adjacent molecules will arrange themselves in such a way that negative portion of that molecule comes near to positive portion of other molecule. It results in the forces of attraction between oppositely charged portions of two adjacent molecules. These attractive forces are called dipole-dipole forces.

$$\overset{\delta+}{H} — \overset{\delta-}{Cl} \cdots\cdots \overset{\delta+}{H} — \overset{\delta-}{Cl} \cdots\cdots \overset{\delta+}{H} — \overset{\delta-}{Cl}$$

Q11. What is a triple covalent bond, explain with an example?

Ans: For answer See Q.3(b)

Q12. What is difference between polar and non-polar covalent bonds, explain with one example of each?

Ans: For answer See Q.10.

Q13. Why does a covalent bond becomes polar?

Ans: The covalent bound becomes polar due to the unequal attraction to the bonded pair of electrons. e.g $\overset{\delta+}{H}—\overset{\delta-}{Cl}$

Q14. What is the relationship between electronegativity and polarity?

Ans: The power of an atom to attract the shared pair of electrons towards itself is called electronegativity.

Relationship:

When a covalent bond is formed between two dissimilar atoms which have a reasonable difference of electronegativity between them, the bond is called polar and the phenomenon is called polarity e.g in $\overset{\delta+}{H}—\overset{\delta-}{Cl}$.

Electronegativity difference is responsible for polarity.

Q15. Why does ice float on water?

Ans: For answer see Q.11.

Q16. Give the characteristics properties of ionic compounds.

Ans: For answer see Q.8.

Q17. What characteristics properties do the covalent compound have?

Ans: For answer see Q.9.

Long Answer Questions

Q1. What is an ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms?

Ans: For answer see Q.2(b)

Q2. How can you justify that bond strength in polar covalent compounds is comparable to that of ionic compound?

Ans: Bond strength in polar covalent compounds depends upon the difference in the electronegativities of bonded atoms. Greater the difference of electronegativities stronger will be the bond. This difference of electronegativity causes the ionic character in covalent compounds. If electronegativity difference is more than 1.7 the bond will be ionic. Hence the strength in polar covalent compounds is comparable to that of ionic compounds.

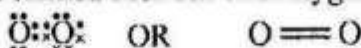
Q3. What type of covalent bonds are formed between hydrogen, oxygen and nitrogen? Explain their bonding with dot and cross model.

Ans: A single covalent bond is formed between hydrogen atoms.

H ×• H OR H — H

Oxygen:

A double covalent bound is formed between two oxygen atoms.

Ö::Ö OR O = O

Q4. How does a covalent bond develop ionic character in it? Explain.

Ans: A double covalent bond is formed between two oxygen atoms.

Pauling suggests that the difference of electronegativity between two bonded atoms tells us the percentage of ionic character is any covalent bond. If the electronegativity difference between two bonded atoms is 1.7 or more than, the bond is ionic. Greater the difference of electronegativity, greater the percentage of ionic character.

Q5. Explain the types of covalent bonds with at least one example of each type.

Ans: For answer see Q.3.

Q6. How is a coordinate covalent bond formed? Explain with examples.

Ans: For answer see Q.4.

Q7. What is metallic bond? Explain the metallic bonding with the help of a diagram.

Ans: For answer see Q.5(b).

Q8. Define hydrogen bonding. Explain how these forces affect the physical properties of compounds.

Ans: Hydrogen bonding exists between the hydrogen atom of one molecule and highly electronegative atom of other molecule.

Hydrogen bonding affects the physical properties of the compounds. Due to this boiling points of the compounds are affected greatly. e.g boiling point of water (100°C) is higher than that of alcohol (78°C). Because of the stronger hydrogen bonding in water.

Q9. What are intermolecular forces? Compare these forces with chemical bond forces with reference to HCl molecule?

Ans: For answer see Q.6.

Q10. What is a chemical bond and why do atoms form a chemical bond?

Ans: For answer see Q.1.

Q11. What is octet rule? Why do atoms always struggle to attain the nearest noble gas electronic configuration?

Ans: For answer see Q.1.

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

4.1 Why do Atoms form Chemical Bonds?

4.2 Chemical Bonds

☆ Tick the correct answer.

1. **Octet rule is:** (FBD. GI)
   (A) Description of eight electron (B) Picture of electronic configuration
   (C) Pattern of electron configuration (D) Attaining of eight electrons
2. **Noble gases are stable because:** (MLN. GII)
   (A) Their valence shell is complete (B) Their valence shell is half filled
   (C) There is no electron in their valence shell
   (D) There are three electrons in their valence shell
3. **Which one of the following is a metal?** (GRW. GI)
   (A) H (B) C (C) N (D) Mg
4. **Which type of forces are dominant during chemical bond formation:** (LHR. GII)
   (A) Repulsive forces (B) Attractive forces
   (C) Vander waal forces (D) Hydrogen bonding

Answers

1. Attaining of eight electrons
2. Their valence shell is complete
3. Mg
4. Attractive forces

☆ Give short answer to the following questions.

1. **How do atoms follow octet rule?** (LHR. GI, SWL. GII, FBD. GII)

Ans. The atoms having 1 to 3 electrons lose electrons, while those having 5,6 or 7 electrons in their valence shell, gain the electrons to complete their octet. The ability of atom to get eight electrons in valence shell is called octet rule.

2. **Define octet rule.** (GRW. GI, MLN. GII, SWL. GI, RWP. GII, BWP. GI, SGD. GI)

Ans. The process of gaining of eight (8) electrons in valence shell is called octet rule.

3. **What is the difference between duplet and octet rule?** (FBD. GI, DGK. GII)

Ans. **Difference between octet and duplet rule:** Gaining of two electrons in valence shell is called duplet rule while gaining of eight electrons in valence shell is known as octet rule.

4. **Why do Atoms react with each other?** (BWP. GII, MLN. GII)

Ans. To become stable and to attain the electronic configuration of noble gases, atoms

react with each other.

5. **Why Noble Gases are not reactive?** (MLN. GII)

Ans. In valence shells of noble gases, their are 2 or 8 electrons. It means noble gases have complete their valence shells. There is no space for more electrons in them, that's why they neither gain nor lose electrons, so these are non reactive.

6. **Define chemical bond. Write down the names of its types.** (RWP. GI, RWP. GII)

Ans. A force between atoms, that hold them in a molecule.

Types of Bond:

i. Ionic bond
ii. covalent bond
iii. Dative covalent bond
iv. Metallic bond

4.3 Types of Chemical Bond

☆ **Tick the correct answer.**

1. **A covalent bond which is formed between two similar atoms is called:** (LHR. GI, MLN GII)
   (A) Metallic bond (B) Non-polar covalent bond
   (C) Polar covalent bond (D) Dative covalent bond
2. **Which one of the following is an electron deficient molecule?** (LHR. GII, SWL. GII, SGD. GII)
   (A) NH_3 (B) BF_3 (C) N_2 (D) O_2
3. **How many covalent bonds does C_2H_2 molecule have?** (LHR. GII, FBD. GII)
   (A) 5 (B) 4 (C) 3 (D) 2
4. **The total number of sharing electrons in N_2 gas is:** (GRW. GI)
   (A) 2 (B) 4 (C) 6 (D) 8
5. **The atom providing lone pair of electrons in dative covalent bond is called:** (GRW. GI)
   (A) acceptor (B) donor (C) electronegative (D) ionic bond
6. **Chlorine has _______ electrons in its outer shell:** (GRW. GII)
   (A) 3 (B) 4 (C) 7 (D) 8
7. **The types of chemical bonds are:** (GRW. GII)
   (A) 1 (B) 2 (C) 3 (D) 4
8. **Which noble gas does not possess 8 electrons in their valence shell?** (FBD. GI)
   (A) He (B) Ne (C) Ar (D) Xe
9. **Bond formed in C_2H_2 is:** (FBD. GII)
   (A) Single (B) Double (C) Triple (D) Metallic
10. **___________ pair has Polar Covalent bonds.** (MLN. GI)
    (A) O_2 and Cl_2 (B) H_2O and N_2 (C) H_2O and C_2H_2 (D) H_2O and HCl
11. **Which one of the following compound is non directional in its bonding?** (SWL. GI)
    (A) CH_4 (B) KBr (C) CO_2 (D) H_2O

12. **Triple covalent bond involves how many electrons?** (SGD. GI, BWP. GII, LHR. GI)
(A) Eight (B) Four (C) Three (D) Six
13. **The bond formed due to mutual sharing of electrons is called:** (SGD. GII)
(A) Metallic bond (B) Ionic bond
(C) Coordinate covalent bond (D) Covalent bond
14. **How many triple covalent bonds are formed in C_2H_2 molecule:** (DGK. GI)
(A) 2 (B) 1 (C) 3 (D) 5
15. **A dative bond is formed between ammonia and boron triflouride, the acceptor atom is:** (DGK. GI)
(A) Flourine (B) Boron (C) Hydrogen (D) Nitrogen
16. **Which type of Covalent bond is present in nitrogen (N_2) molecule:** (LHR. GII)
(A) Single Covalent bond (B) Double Covalent bond
(C) Triple Covalent bond (D) Metallic bond
17. **Metals have generally:** (LHR. GII)
(A) High ionization value (B) Low ionization value
(C) High electron affinity value (D) High electro-negativity value
18. **In triple covalent bond, each bonded atom contributes electrons:** (MLN. GI)
(A) 4 (B) 2 (C) 3 (D) 6
19. **The sodium atom after losing one electron attain the electronic configuration that is:** (MLN. GII)
(A) $1s^2 2s^2 2p^6 3s^2$ (B) $1s^2 2s^2 2p^6$ (C) $1s^2 2s^2 2p^5$ (D) $1s^2 2s^2 2p^4$
20. **Which molecule needs two electrons to complete its outermost shell?** (SWL. GI)
(A) N_2 (B) O_2 (C) NH_3 (D) BF_3
21. **Bond formation between ions is due to:** (SWL. GII)
(A) electron sharing (B) intermolecular forces
(C) electrostatic forces (D) repulsive forces
22. **The difference of electronegativity between two elements is more than 1.7 the bond will be:** (SGD. GI)
(A) Covalent bond (B) Ionic bond (C) Non polar (D) None
23. **Chemical bond formed between two similar atoms is:** (SGD. GII)
(A) polar bond (B) non polar bond
(C) metallic bond (D) dative covalent bond
24. **The number of electrons in the valence shell of Noble gases is:** (RWP. GI)
(A) 8 (B) 7 (C) 6 (D) 17
25. **The formation of ammonium ion $[NH_4]^+$ is due to:** (RWP. GI)
(A) covalent bond (B) ionic bond
(C) metallic bond (D) co-ordinate covalent bond
26. **Noble gases have electrons in their valance shell:** (RWP. GII)
(A) 2 or 8 (B) 2 or 6 (C) 2 or 4 (D) 2 or 10

27. **Covalent bond is found in methane (CH_4).** (RWP. GII)
(A) single (B) double (C) triple (D) dative

28. **A bond pair in covalent molecules usually has:** (DGK. GI)
(A) one electron (B) two electrons (C) three electrons (D) four electrons

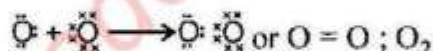
Answers

1. Non-polar covalent bond 2. BF_3 3. 5
4. 6 5. donor 6. 7 7. 4
8. He 9. Triple 10. H_2O and HCl 11. KBr
12. Six 13. Covalent bond 14. 1 15. Boron
16. Triple Covalent bond 17. High ionization value
18. 3 19. $1s^2\ 2s^2\ 2p^6$ 20. BF_3 21. electrostatic forces
22. Ionic bond 23. non polar bond 24. 8 25. co-ordinate covalent bond
26. 2 or 8 27. single 28. two electrons

☆ **Give short answer to the following questions.**

1. **Explain double covalent bond with the help of an example.** (LHR. GI, DGK. GII)

Ans. A double covalent bond is formed by the mutual sharing of two pairs of electrons. It is represented by double lines. This type of double covalent bond is formed in Oxygen (O_2) gas.

$$\ddot{O} + :\ddot{O} \longrightarrow \ddot{O}::\ddot{O} \text{ or } O = O\ ;\ O_2$$

2. **Why does sodium make ionic bond with chlorine?** (LHR. GII)

Ans. Sodium is an electropositive element. It's ionization energy is very low. It has tendency to lose electron and become Na^+. while Cl is an electronegative element. It has high electron affinity, so it attract the electron and become Cl^-. They complete their valence shell with 8 electrons. Both ions stable themselves by electrostatic force of attraction.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

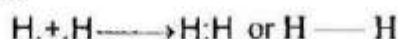
3. **Which type of covalent bond is formed in N_2 gas?** (LHR. GII, SGD. GII)

Ans. There is triple covalent bond in nitrogen gas.

$:N \equiv N:$

4. **Define non-polar covalent bond and give an example.** (GRW. GI & II, MLN. GI)

Ans. When a covalent bond is formed between two similar atoms, the shared pair of electrons are attracted by both the atoms equally, such type of bond is called non-polar covalent bond.

$$H. + .H \longrightarrow H:H \text{ or } H - H$$

5. **What is meant by triple covalent bond? Explain it with an example.** (GRW. GII, RWP. GI)

Ans. A type of covalent bond formed by mutual sharing of six electrons. It is represented by three short lines between two atoms. e.g Nitrogen molecules have

triple covalent bond. :N≡N:

6. Draw the lewis structure diagram of $C\ell_2$. (FBD. GI)

Ans. Lewis Structure diagram: $:\ddot{\underset{..}{Cl}}\cdot\times\ddot{\underset{\times\times}{Cl}}\times$

7. Write the electronic configuration of $C\ell^-$ ion. (SWL. GI)

Ans. $Cl^- = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

8. What is metallic bond? (SGD. GII)

Ans. A bond formed due to mobile electrons between metallic atoms (positive charged ions).

9. Why does Chlorine accept an electron and attain -1 charge? (RWP. GII, GRW. GII)

Ans. Chlorine has seven electrons in its valence shell, to complete its last orbital, it required only one electron. So to make it stable, either it has to give up seven electrons or gain one electron. As it is easy to gain one electron, instead of releasing seven electrons, it gain one electron and attain -1 charge.

10. Define Ionic bond with an example. (DGK. GI)

Ans. The type of chemical bond in which electron is transferred from one atom to another, called ionic bond.

11. How many bond pairs of electrons are found in NH_3 Molecule? (BWP. GI)

Ans. There are three bonded pair of electrons in the molecules of NH_3.

12. Why a Covalent bond becomes polar? (LHR. GII, GRW. GII)

Ans. When there is difference of electro-negativity between two bonded atom, then there is unequal attraction of bond pair between these atoms, resulting in formation of polar covalent bond.

13. What is difference between ionic bond and covalent bond? (SWL. GI, FBD. GI)

Ans. Ionic bond: A type of chemical bond, which is formed by complete transfer of electron from one atom to another.

$2Na + C\ell_2 \longrightarrow 2NaC\ell$

Covalent bond: A type of bond formed by sharing of electrons between two atoms.

$H. + \times H \longrightarrow H.\times H$

14. Point out the type of covalent bond in the following molecules. CH_4, C_2H_4, H_2, N_2 and O_2 (SWL. GI)

Ans. H_2, CH_4 = Single covalent bond

O_2, CH_4 = Double covalent bond

N_2 = Triple covalent bond

15. What is difference between bond pair and lone pair of electrons? Explain with an example. (DGK. GI, LHR. GI, SWL. GII)

Ans. Bond pair: A pair of electron in a molecule, that take part in bonding.

Lone pair: Non-bonding electron pair present on atom of molecule is called lone pair.

16. Why does oxygen molecule not form a polar covalent bond? (RWP. GI)

Ans. In oxygen molecule, there is equal sharing of bonded pair between two atoms. Due to this polar covalent bond is not formed.

17. Differentiate between polar and non-polar compounds. (RWP. GI)

Ans. **Polar compounds:** Those compounds, which are formed by combining of different type of atoms having electronegativity difference.

e.g $H^{+} \longrightarrow Cl^{-}$

Non-polar compound:

Those compound, which are formed by combinng same type of atoms.

e.g $H. + \times H \longrightarrow H\dot{\times}H$

4.4 Intermolecular Forces
4.5 Nature of Bonding and Properties

☆ Tick the correct answer.

1. **The melting point of sodium chloride is:** (LHR. GI, SGD. GI, FBD. GII, MLN. GII)
   (A) 600°C (B) 750°C (C) 800°C (D) 1000°C
2. **Which one of the following compound is not soluble in water.** (SWL. GI, DGK. GI)
   (A) $MgCl_2$ (B) $NaCl$ (C) KBr (D) C_6H_6
3. **Ice floats on water because:** (RWP. GI)
   (A) Ice is denser than water (B) Ice is crystalline in nature
   (C) Water is denser than ice (D) Water molecules move randomly
4. **The weakest force among the atoms is:** (RWP. GII)
   (A) Ionic force (B) Metallic force (C) Intermolecular force (D) Covalent force
5. **Which type of force is present in hydrogen bonding:** (LHR. GI)
   (A) Inter molecular force (B) Ionic force
   (C) Covalent force (D) Metallic force
6. **Which type of bond is present in H - F molecule?** (DGK GII)
   (A) ionic (B) non polar (C) polar covalent (D) coordinate
7. **Hydrogen bonding has:** (BWP. GII)
   (A) intermolecular force (B) ionic force (C) covalent force (D) metallic force
8. **The example of ionic compound is:** (SWL. GII, BWP. GII, GRW. GI)
   (A) $NaCl$ (B) H_2 (C) HCl (D) O_2
9. **The boiling point of sodium chloride is:** (GRW. GI)
   (A) 1413° (B) 1513° (C) 1613° (D) 1713°
10. **Which one of the following is a polar molecule?** (GRW. GII)
    (A) CH_4 (B) H_2 (C) Cl_2 (D) H_2O
11. **Metals are generally good conductors of electricity because:** (BWP. GI)
    (A) mobile electrons are present (B) metal cations are present

(C) they are hard enough (D) mobile protons are present

12. Ionic character of bond becomes dominant over covalent character when: (BWP. GI)

(A) if electronegativity difference is greater than 1.7
(B) if electronegativity difference is less than 1.7
(C) if electronegativity difference is equal than 1.7
(D) if electronegativity difference is equat to zero

Answers

1. 800°C 2. C_6H_6 3. Ice is denser than water
4. Intermolecular force 5. Inter molecular force 6. polar covalent
7. intermolecular force 8. NaCl 9. 1413° 10. H_2O
11. mobile electrons are present
12. if electronegativity difference is greater than 1.7

☆ Give short answer to the following questions.

1. Explain polar covalent bond with an example. (LHR. GI, FBD. GII, GRW. G I)

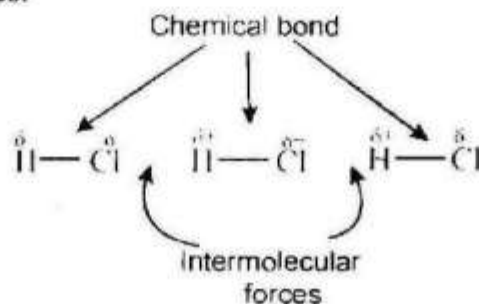
Ans. Covalent bond which is formed between two dissimilar atoms having a reasonable difference of electronegativities between bonded atoms is called polar covalent bond. The electronegativity difference of hydrogen and chlorine is 1.0. As chlorine has more electronegativity than hydrogen. That is why it attract the common pair of electron toward it with greater force. So due to this difference of electronegativity there is partial negative charge on chlorine and partial positive charge on hydrogen. A polarity is being developed, that is why it is known as polar covalent bond.

$$H\cdot + \cdot\ddot{\underset{..}{Cl}}: \longrightarrow H^{\delta+} \cdot\cdot \ddot{\underset{..}{Cl}}:^{\delta-}$$

2. Define hydrogen bonding. (GRW. GI, SGD. GI, MLN. GI, DGK. GII)

Ans. Hydrogen bonding: The electrostatic force of attraction between highly electronegative small atoms like oxygen nitrogen, fluorine and partial positive hydrogen atom is called hydrogen bonding. It is represented by dotted line (….).

3. Define the inter molecular force of attraction. (FBD. GI, MLN. GII)

Ans. The forces of attraction present between the molecules of a compound are called intermolecular forces.

4. Why does ice float on water? (FBD. GII, MLN. GI, BWP. GI)

Ans. When water is cooled, it become freeze due to hydrogen bonding. The distance between water molecules increases at specific distance, resulting in the density of ice becomes less than that of water. That's the reason why ice floats on water.

5. What is difference between polar and non-polar covalent bonds? (SWL. GII, SGD. GII, RWP. GII)

Ans.

Polar covalent bond	Non-polar covalent bond
A type of covalent bond formed between two different atoms, having electro negativity difference. $H^{+\delta} - Cl^{-\delta}$	A type of covalent bond formed between same type of atoms called non-polar covalent bond. $H. + .H \longrightarrow H:H$

6. What do you mean by malleability? (FBD. GII)

Ans. Malleability is a property by virtue of which it can be drawn into sheets.

7. Metals are good conductors of electricity, why? (SWL. GI, MLN. GII, RWP. GII)

Ans. Metals are good conductor of electricity, because they have mobile free electrons. When electric field is applied across the metals, the free electrons coming toward positive pole, got new space. This pressure of electron force the free electrons to move. In this way current is passed through metal.

8. Write any two properties of ionic compounds. (SGD. GI)

Ans. Following are the properties of ionic compounds:

i. Ionic compounds are crystalline solids.

ii. In solid form there is trace of electrical conductance of ionic compound. But in solution or melted form they are good conductors of electricity.

9. Ionic bond is stronger than covalent bond explain. (SGD. G II)

Ans. The ions of ionic compound have strong electrostatic force. Due to which, they stable strongly on their positions. That is why ionic bond is more stronger than covalent bond.

10. Ionic compounds are solids. Justify. (RWP. GII)

Ans. Ionic compounds are solids because ionic compound, have very strong ionic bonds, combine ions to give solids.

11. Write down any four properties of Metals. (MLN. GI)

Ans. 1. All metals are solids except mercury.

2. They have high melting and boiling point.
3. They are good conductor of electricity.
4. They have metallic bonding.

PHYSICAL STATES OF MATTER

Major Concepts:

Gaseous State
5.1 Typical properties 5.2 Laws related to gases
Liquid State: 5.3 Typical Properties
Solid State: 5.4 Typical Properties
5.5 Types of Solids 5.6 Allotropy

Time allocation

Teaching periods	16
Assessment periods	04
Weightage	10%

Students Learning Outcomes:

Students will be able to:

- Effect on the volume of a gas by a change in the a. pressure b. temperature.
- Compare the physical states of matter with regard to intermolecular forces present between them.
- Account for pressure-volume changes in a gas using Boyle's Law.
- Account for temperature-volume change in a gas using Charles' Law.
- Explain the properties of gasses (diffusion, effusion and pressure).
- Explain the properties of liquids like evaporation, vapour pressure, boiling point.
- Explain the effect of temperature and external pressure on vapour pressure and boiling point.
- Describe physical properties of solids (melting and boiling points).
- Differentiate between amorphous and crystalline solids.
- Explain the allotropic forms of solids.

5.1 TYPICAL PROPERTIES OF GASEOUS STATE

Q.1 What is meant by matter? Name the three states of matter and compare their properties.

Ans. Matter: Any thing that has mass and occupies space in called matter e.g book water, air etc.

States of matter:

Matter exists in three physical states. (i) Solid (ii) Liquid (iii) Gas

Solids	Liquids	Gases
1. Solids have fixed shape and fixed volume.	Liquids have fixed volume but no fixed shape. It attains the shape of container in which it is kept.	They have no fixed shape and volume.

2. The particles of solids are fixed in their mean positions and only vibrate about their mean positions.	The particles of liquid are much farther a part than those of solids.	The particles of gases are freely moving in all possible directions.

Q.2 Explain following typical properties of gases.

(i) Diffusion (ii) Effusion (iii) Pressure (iv) Compressibility (v) Mobility (vi) Density of gases

Ans. (i) Diffusion: The movement of the molecules of a substance from higher concentration to lower concentration is called diffusion.

OR

The spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture.

Rate of diffusion: The rate of diffusion depends upon the molecular mass, lighter gases diffuse rapidly than heavier ones. e.g. H_2 diffuses four times faster than O_2 gas.

(ii) Effusion: The escaping of gas molecules through a tiny hole into a space with lesser pressure. When a tyre gets punctured, air effuses out. Effusion depends upon the molecular masses, lighter gases effuse faster than heavier gases.

(iii) Pressure: Pressure in defined as force per unit area.

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

$$P = F/A$$

Units: The SI unit of pressure is Nm^{-2}. It is also called Pascal (Pa).

One Pascal (Pa) = $1Nm^{-2}$

Measurement of pressure: Barometer is used to measure atmospheric pressure and manometer is used to measure pressure in the laboratory.

Standard atmospheric pressure:

It is the pressure exerted by the atmosphere at sea level. It is defined as.

The pressure exerted by the column of mercury 760 mm height at sea level.

1atm = 760mm of Hg = 760 torr

1mm of Hg = one torr

$101325Nm^{-2}$ = 101325 Pa

(iv) Compressibility:

Gases are highly compressible due to empty spaces between their molecules.

(v) Mobility: Gas molecules are always in state of continuous motion. They can move from one place to another because gas molecules possess very high kinetic energy.

(vi) Density: Density is defined as mass per unit volume i.e.

$$D = \frac{m}{V}$$

Density of gases: Gases have low density than liquids and solids. It is due to light mass

and more volume occupied by gas molecules.

The density of gases increases by cooling because their volume decreases. e.g. at normal atmospheric pressure the density of oxygen gas is $1.4 gdm^{-3}$ at 20°C and $1.5 gdm^{-3}$ at 0°C.

Units: The density of gases is expressed in grams per dm^3 (gdm^{-3}) whereas, the density of liquids and solids are expressed in grams per cm^3. (gcm^{-3}) [solids and liquids are 1000 times denser than gases.]

Test yourself: 5.1:

(i) Why the rate of diffusion of gases is rapid than that of liquids?

Ans. The rate of diffusion of gases is rapid than liquid due to the less attractive forces among the gas molecules than liquids.

(ii) Why are the gases compressible?

Ans. Gases are compressible due to the availability of large spaces among the gas molecules.

(iii) What do you mean by Pascal? How many Pascals are equal to 1 atm?

Ans. Pascal: Pascal is the unit of pressure.

One Pascal $= 1 Nm^{-2}$

101325pa = 1 atm.

(iv) Why the density of a gas increases on cooling?

Ans. Density of gases increases by cooling because their volume decreases.

(v) Why is the density of gas measured in $g\ dm^{-3}$ while that of a liquid in $g\ cm^{-3}$?

Ans. Density of gases is less than that of liquid hence the density of gases is measured in gdm^{-3} and that of liquid in gcm^{-3}.

(vi) Convert the following: (a) 70 cm Hg to atm (b) 3.5 atm to torr (c) 1.5 atm to Pa

Ans. (a) 70cm Hg to atm.

76cm Hg = 1 atm

$$70\text{cm Hg} = \frac{1}{76} \times 70 = 0.92 \text{ atm.}$$

(b) 3.5 atm to torr

1 atm = 760 torr

3.5 atm = 760×3.5

= 2660 torr

(c) 1.5 atm to Pa

1 atm = 101325 Pa

1.5 atm = 101325×1.5

= 151987.5 Pa.

5.2 LAWS RELATED TO GASES

Q.3 What is meant by Boyle's law? Give its experimental verification.

Ans. Boyle's Law: This law was put forward by Robert Boyle in 1662. This law states that;

"The volume of a given mass of a gas is inversely proportional to its pressure at constant temperature.

Mathematical formula: It can be written as;

$$\text{Volume} \propto \frac{1}{\text{pressure}}$$

$$V \propto \frac{1}{P}$$

Or $$V = k.\frac{1}{P}$$

Or $$PV = k$$

Rober Boyle (1627-1691) was natural philosopher, chemist, physicist and inventor. He is famous for 'Boyle's law of gases'.

$$\therefore \ PV = \text{Constant}$$

Where "k" is the constant of proportionality. Boyle's law can also be stated as;

"The product of pressure and volume of a fixed mass of a gas is constant at a constant temperature. If

$$P_1V_1 = k$$

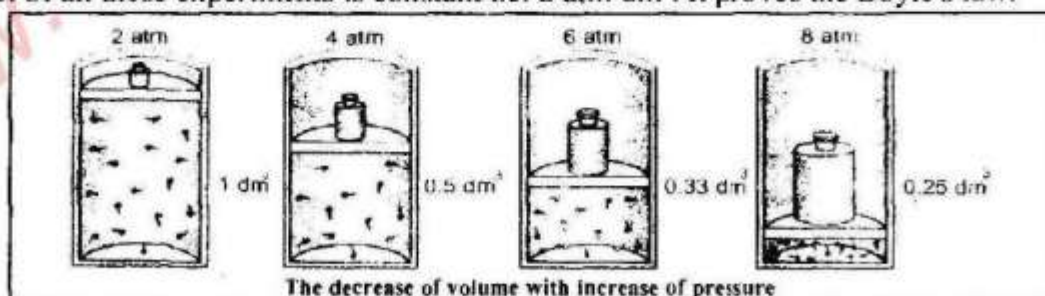
Then $$P_2V_2 = k$$

$$P_1V_1 = P_2V_2$$

Where P_1 = initial pressure, P_2 = Final pressure, V_1 = initial volume, V_2 = Final volume.

Experimental verification of Boyl's law: Take some mass of a gas in a cylinder having a movable piston and observe the effect of increase of pressure on its volume. When the pressure of 2 atmosphere (atm) is applied, the volume of the gas reads as 1 dm^3. When pressure is increased equivalent to 4 atm, the volume of the gas reduces to $0.5dm^3$. Again when pressure is increased three times i.e. 6 atm, the volume reduces to $0.33dm^3$. Similarly, when pressure is increased up to 8 atm on the piston, volume of the gas decreases to $0.25dm^3$.

When we calculate the product of volume and pressure for this experiment, the product of all these experiments is constant i.e. 2 atm dm^3. It proves the Boyle's law.

The decrease of volume with increase of pressure

$P_1V_1 = 2atm \times 1dm^3 = 2\ atm\ dm^3$

$P_2V_2 = 4atm \times 0.5dm^3 = 2atm\ dm^3$

$P_3V_3 = 6atm \times 0.33dm^3 = 2atm\ dm^3$

$P_4V_4 = 8\ atm \times 0.25dm^3 = 2atm\ dm^3$

Test yourself 5.2:

(i) **Is the Boyle's law applicable to liquids?**

Ans. No, it is only applicable to gases.

(ii) **Is the Boyle's law valid at very high temperature?**

Ans. Yes it is applicable at high (constant) temperature.

(iii) **What will happen if the pressure on a sample of gas is raised three times and its temperature is kept constant?**

Ans. The volume of the gas will decrease three times.

Do you know?

In Which units blood pressure is measured?

Blood pressure is measured using a pressure gauge. It may be a mercury manometer or some other device. Blood pressure is reported by two values, such as 120/80, which is a normal blood pressure. The first measurement shows the maximum pressure when the heart is pumping. It is called systolic pressure. When the heart is in resting position, pressure decreases and it is the second value called diastolic. Both of these pressures are measured in torr units. Hypertension is because of high blood pressure due to tension and worries in daily life. The usual criterion for hypertension is a blood pressure greater than 140/90. Hypertension raises the level of stress on the heart and on the blood vessels. This stress increases the susceptibility of heart attacks and strokes.

Example 5.1

A gas with volume $350cm^3$ has a pressure of 650mm of Hg. If its pressure is reduced to 325mm of Hg, calculate what will be its new volume?

Data

$V_1 = 350cm^3$

$P_1 = 650mm$ of Hg

$P_2 = 325mm$ of Hg

$V_2 = ?$

Solution:

By using the equation of Boyle's Law

$$P_1V_1 = P_2V_2 \quad \text{or} \quad V_2 = \frac{P_1V_1}{P_2}$$

By putting the values;

$$V_2 = \frac{650\times 350}{325} = 700cm^3$$

Thus volume of the gas is doubled by reducing its pressure to half.

Example 5.2

$785cm^3$ of a gas was enclosed in a container under a pressure of 600mm Hg. If volume is reduced to $350cm^3$, What will be the pressure?

Data

$V_1 = 785cm^3$

$P_1 = 600\text{mm of Hg}$

$V_2 = 350\text{cm}^3$

$P_2 = ?$

Solution:

By using the Boyle's equation

$$P_1V_1 = P_2V_2 \quad \text{or} \quad P_2 = \frac{P_1V_1}{V_2}$$

By putting the values

$$P_2 = \frac{785 \times 600}{350} = 1345.7\text{mm of Hg}$$

$$\text{Or} \quad P_2 = \frac{1345.7}{760} = 1.77\text{atm}$$

Thus pressure is increased by decreasing volume.

Q.4 What is meant by Charles's law? Give its experimental verification.

Ans. Charles's law: This law was put forward by a French scientist J. Charles in 1787. This law states that

"The volume of a given mass of a gas is directly proportional to the absolute temperature at constant pressure."

Mathematical formula: Volume $\propto$ Temperature

$$V \propto T \quad \text{or} \quad V = kT \quad \text{or} \quad \frac{V}{T} = k$$

Where "k" is the constant of proportionality

If the temperature is increased from T_1 to T_2, the volume changes from V_1 to V_2.

Mathematically it can be written as:

$$\frac{V_1}{T_1} = k$$

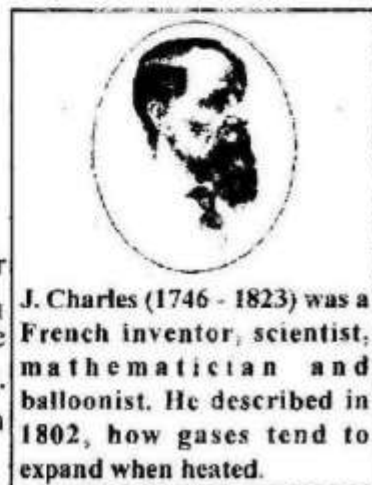
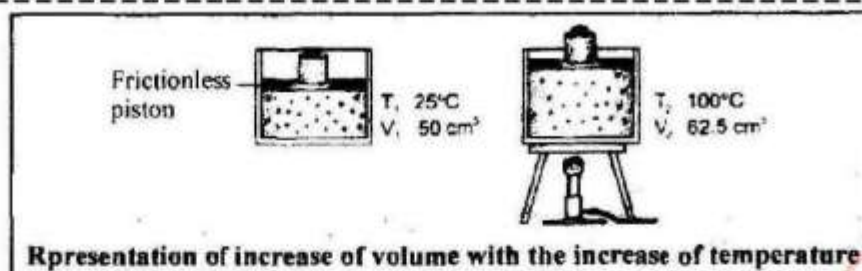
$$\text{and} \quad \frac{V_2}{T_2} = k$$

$$\therefore \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

J. Charles (1746 - 1823) was a French inventor, scientist, mathematician and balloonist. He described in 1802, how gases tend to expand when heated.

Experimental verification of charles's law:

Take a certain amount of gas enclosed in a cylinder having a movable piston. If the initial volume of the gas V_1 is 50 cm³ and initial temperature T_1 is 25°C, on heating the cylinder up to 100°C, its new volume V_2 is about 62.5cm³. The increase in temperature, increases the volume as shown in figure below.

Rpresentation of increase of volume with the increase of temperature

Q.5 (a) Write a note on absolute temperature.

(b) Describe the role of intermolecular forces in the physical states of matter.

Ans. Absolute temperature scale:

Lord Kelvin introduced absolute temperature scale or Kelvin scale.

Conversion of Kelvin and Centigrade (Celsius) temperature.

This scale of temperature starts from 0K or -273.15C°, which is given the name of absolute zero.

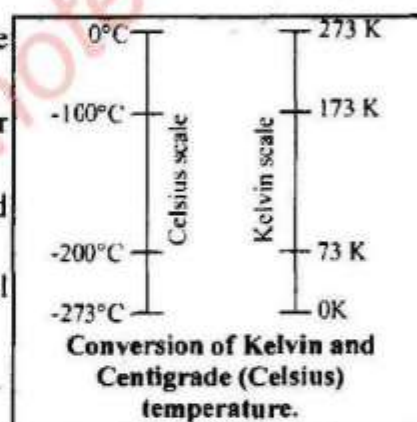
[It is the temperature at which an ideal gas would have zero volume.]

Both Kelvin and Centigrade scale have equal degree range.

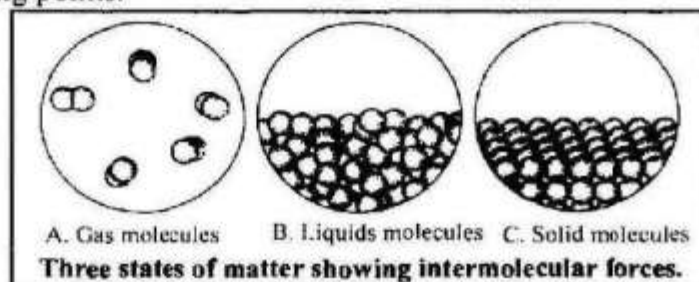
0K = -273°C

and 273K = 0°C

K = C° + 273

C° = K – 273

(b) Role of Intermolecular Forces in the Physical States of Matter:

Matter exists in three physical states, gas, liquid and solid. In the gaseous state, the molecules are far apart from each other. Therefore, intermolecular forces are very weak in them. But in the liquid and solid states intermolecular forces play a very important role on their properties. In the liquid state molecules are much closer to each other as compared to gases as shown in figure. As a result liquid molecules develop stronger intermolecular forces, which affect their physical properties like diffusion, evaporation, vapour pressure and boiling point. Compounds having stronger intermolecular forces have higher boiling points.

Three states of matter showing intermolecular forces.

The intermolecular forces become so dominant in solid state that the molecules look motionless. They arrange in a regular pattern therefore they are denser than molecules of liquids.

Example 5.3

A sample of oxygen gas has a volume of 250cm³ at -30°C. If gas is allowed to expand up to 700cm³ at constant pressure, find out its final temperature.

Data

$V_1 = 250cm^3$

$T_1 = -30°C = (-30 + 273) = 243K$

$V_2 = 700cm^3$

$T_2 = ?$

Solution:

By using the equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad T_2 = \frac{V_2T_1}{V_1}$$

By putting the value in equation.

$$T_2 = \frac{700 \times 243}{250} = 680.4K$$

Thus expansion is caused due to increasing temperature.

Example 5.4

A sample of hydrogen gas occupies a volume 160cm³ at 30°C. If its temperature is raised to 100°C, calculate what will be its volume if the pressure remains constant.

Data:

$V_1 = 160cm^3$

$T_1 = 30°C = 303K$ (as $0°C = 273K$)

$V_2 = 100°C = 373K$

$V_2 = ?$

Solution:

By using the equation of Charles Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{or} \quad V_2 = \frac{V_1T_2}{T_1}$$

By putting the values in above equation: $V_2 = \dfrac{160 \times 373}{303} = 196.9cm^3$

Thus volume of the gas has increased by raising the temperature.

Test yourself 5.3:

(i) Which variables are kept constant in Charles's law?

Ans. Pressure is kept constant in charles's law.

(ii) Why volume of a gas decreases with increase of pressure?

Ans. The volume of gas decreases by increasing pressure because there are large spaces present among the gas molecules, when the pressure is applied, the molecules become closer to each other and their volume will decrease.

(iii) What is absolute zero?

Ans. It is the temperature at which an ideal gas would have zero volume. Its value is -273.15°C.

(iv) Does Kelvin scale show a negative temperature?

Ans. No Kelvin scale can not show negative temperature. The minimum temperature at this scale is 0 K which is equal to -273. The temperature below this is not possible.

(v) When a gas is allowed to expand, what will be its effect on its temperature?

Ans. At constant pressure, if volume is increased, the temperature will increased.

(vi) Can you cool a gas by increasing its volume?

Ans. No, the gas can not be cooled by increasing volume.

Do you know?

In which units' body temperature is measured?

Body temperature is measured in Fahrenheit scales. Normal body temperature is 98.6 °F, it is equivalent to 37 °C. This temperature is close to average normal atmospheric temperature. In winter atmospheric temperature falls lower than that of our body temperature.

According to principle of heat flow, heat flows out from our body and we feel cold. To control this outward flow of heat, we wear black and warm clothes. To maintain body temperature we use dry fruits, tea, coffee and meats etc.

5.3 TYPICAL PROPERTIES OF LIQUID STATE

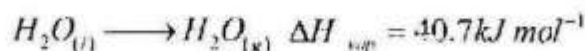
Q.6(a) What is meant by evaporation? Explain how does the process of evaporation occur?

(b) Evaporation is a cooling process. Justify.

(c) Discuss the factors on which evaporation depends?

Ans. (a) Evaporation: The continuous escape of the molecules of a liquid from its surface is called evaporation and it occurs at any temperature.

Evaporation is reverse of condensation. It is an endothermic process. e.g.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)} \quad \Delta H_{vap} = 40.7 kJ\ mol^{-1}$$

How does evaporation occur: In the liquid state, molecules are in a continuous state of motion. They possess kinetic energy but all the molecules do not have same kinetic energy. Majority of the molecules have average kinetic energy and a few have more than average kinetic energy. The molecules having more than average kinetic energy overcome the attractive forces among the molecules and escape from the surface. It is called as evaporation. Evaporation is a continuous process taking place at all temperatures.

Rate of evaporation: The rate of evaporation is directly proportional to temperature. It

increases with the increase in temperature because of increase in kinetic energy of the molecules.

(b) Evaporation is a cooling process: Evaporation is a cooling process. When the high kinetic energy molecules vapourize, the temperature of remaining molecules falls down. To compensate this deficiency of energy, the molecules of liquid absorb energy from the surroundings. As a result the temperature of surroundings decreases and we feel cooling effect. For example, when we put a drop of alcohol on palm, the alcohol evaporates and we feel cooling effect.

(C) Factors on which evaporation depends:

Evaporation depends upon following factors.

(i) Surface area: Evaporation is a surface phenomenon. Greater is surface area greater is evaporation and vice verse. For example, sometimes a saucer is used if tea is to be cooled quickly. This is because evaporation from the larger surface area of saucer is more than that from the smaller surface area of a tea cup.

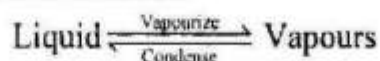
(ii) Temperature: At high temperature, rate of evaporation is high because at high temperature kinetic energy of the molecules increases so high that they overcome the intermolecular forces and evaporate rapidly. For example, water level in a container with hot water decreases earlier than that of a container with cold water. This is because the hot water evaporates earlier than the cold water.

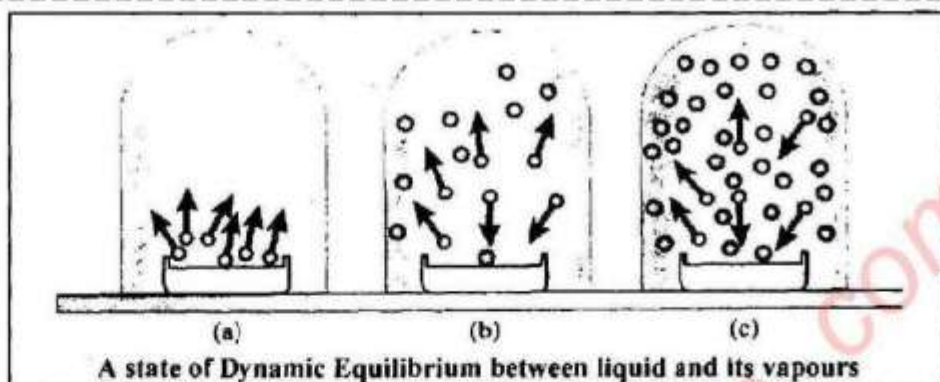
(iii) Intermolecular forces: If intermolecular forces are stronger, molecules face difficulty in evaporation. For example, water has stronger intermolecular forces than alcohol, therefore, alcohol evaporates faster than water.

Q.7 (a) What is meant by vapour pressure? Explain it.

(b) Discuss the factors on which the vapour pressure of a liquid depends?

Ans. (a) Vapour pressure: The pressure exerted by the vapours of a liquid at equilibrium with the liquid at a particular temperature is called vapour pressure of a liquid. The equilibrium is a state when rate of vapourization and rate of condensation is equal to each other but in opposite direction.

$$\text{Liquid} \underset{\text{Condense}}{\overset{\text{Vapourize}}{\rightleftharpoons}} \text{Vapours}$$

Explanation: From the open surface of a liquid, molecules, evaporate and mix up with the air but when we close a system, evaporated molecules start gathering over the liquid surface. Initially the vapours condense slowly to return to liquid. After sometime condensation process increases and a stage reaches when the rate of evaporation becomes equal to rate of condensation. At that stage the number of molecules evaporating will be equal to the number of molecules coming back (condensing) to liquid. This state is called dynamic equilibrium as a shown in figure.

A state of Dynamic Equilibrium between liquid and its vapours

(b) Factors on which the vapour pressure of a liquid depends:

Vapour pressure of a liquid depends upon the following factors.

(i) Nature of liquid: Vapour pressure depends upon the nature of liquid. Polar liquids have low vapour pressure than non-polar liquids at the same temperature. This is because of strong intermolecular forces between the polar molecules of liquids. For example, water has less vapour pressure than that of alcohol at same temperature.

(ii) Size of molecules: Small sized molecules can easily evaporate than big sized molecules hence, small sized molecules exert more vapour pressure. For example, hexane (C_6H_{14}) is a small sized molecule as compared to decane ($C_{10}H_{22}$). C_6H_{14} evaporates rapidly and exerts more pressure than $C_{10}H_{22}$.

(iii) Temperature: At high temperature, vapour pressure is higher than at low temperature. At elevated temperature, the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

Relationship of Vapour Pressure of Water with Temperature

Temp °C	Vapour Pressure mmHg	Temp °C	Vapour Pressure mmHg
0	4.58	60	149.4
20	17.5	80	355.1
40	55.3	100	760.0

Q.8 Define and explain the boiling point. Discuss the factors on which boiling point depends.

Ans. Boiling point: Boiling point is defined as the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.

Explanation: When a liquid is heated, its molecules gain energy. The number of molecules which have more than average kinetic energy increases. More and more molecules become energetic enough to overcome the intermolecular forces. Due to this, rate of evaporation increases that results in increase of vapour pressure until a stage reaches where the vapour pressure of a liquid becomes equal to atmospheric pressure. At this stage the liquid starts boiling.

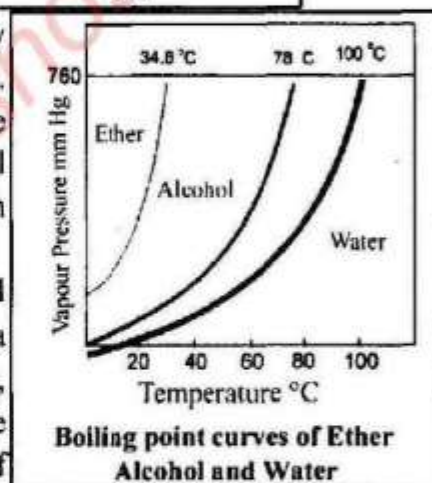
Factors on which boiling point depends: The boiling point of the liquid depends upon the following factor.

(i) Nature of Liquid: The polar liquids have high boiling points than that of non-polar liquids because polar liquids have difficulty in evaporation. Boiling points of a few liquids are given in the table.

Sr. No.	Liquid	Boiling Point °C
1.	Diethyl ether	34.6
2.	Ethyl ether	78
3.	Water	100
4.	n – octane	126
5.	Acetic acid	118

(ii) Intermolecular forces: Intermolecular forces play a very important role on the boiling point of liquids. Substances having stronger intermolecular forces have high boiling points, because such liquids attain a level of vapour pressure equal to external pressure at high temperature. It is given in figure.

Boiling point curves of Ether Alcohol and Water

(iii) External pressure: Boiling points of a liquid depends upon external pressure. Boiling point of a liquid is controlled by external pressure in such a way, that it can be increased by increasing external pressure and vice versa. This principle is used in the working of 'pressure cooker.

Q9. Define freezing point. Write down the freezing points of some common liquids.

Ans. Freezing point:

Freezing point of a liquid is that temperature at which vapour pressure of liquid phase is equal to the vapour pressure of the solid phase. At this temperature liquid and solid coexist in dynamic equilibrium with one another.

Freezing Points of Common Liquids

Sr. No.	Liquid	Freezing Point °C
1.	Diethyl ether	–116
2.	Ethyl alcohol	–115
3.	Water	0.0
4.	n – octane	–5.7
5.	Acetic acid	16.6

Q10. (a) What is meant by diffusion, explain the diffusion in liquid with an example?

(b) Describe the factors which affect the rate of diffusion of liquid.

Ans. Diffusion:

The movement of the molecules of a substance from higher to lower concentration is called diffusion.

Diffusion in liquids

Diffusion in liquids: The liquid molecules are always in a state of continuous motion. They move from higher concentration to lower concentration. They mix up with the molecules of other liquids, so that they form a homogenous mixture.

Example: When a few drops of ink are added in a beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place. Liquids diffuse like gases but the rate of diffusion of liquid is very slow.

(b) Factors affecting the rate of diffusion of liquid:

The diffusion of liquid depends upon the following factors.

(i) Intermolecular forces: Liquids having weak intermolecular forces diffuse faster than those having strong intermolecular forces.

(ii) Size of molecules: Big sized molecules diffuse slowly. For example, honey diffuses slowly in water than that of alcohol in water.

(iii) Shapes of molecules: Regular shaped molecules diffuse faster than irregular shaped molecules because they can easily slip over and move faster.

(iv) Temperature: Diffusion increases by increasing temperature because at high temperature the intermolecular forces are weak.

Q11. What is meant by density? Describe the density of liquid.

Ans. Density: Density is defined as the mass per unit volume.

$$D = m/V$$

Density of liquids: The density of liquid depends upon its mass and volume. Liquids are denser than gases because molecules of liquid are closely packed and the spaces between their molecules are negligible. Liquid molecules have strong intermolecular forces hence they cannot expand freely and have a fixed volume. Like gases, they cannot occupy all the available volume of the container that is the reason why densities of liquids are high. For example: density of water is $1.0 gcm^{-3}$ while that of air is $0.001 g\ cm^{-3}$. That is the reason why drops of rain fall downward. The densities of liquids also vary. You can observe kerosene oil floats over water while honey settles down in the water.

Test yourself 5.4

(i) Why does evaporation increase with the increase of temperature?

Ans. At high temperature the rate of evaporation increases because at high temperature the kinetic energy of the molecules increases, they overcome the intermolecular forces and evaporate easily.

(ii) What do you mean by condensation?

Ans. The process of converting gases into liquid state is called condensation.

(iii) Why is vapour pressure higher at high temperature?

Ans. At high temperature the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

(iv) Why is the boiling point of water higher than that of alcohol?

Ans. Boiling point of water is higher than alcohol due to the presence of strong hydrogen bonding in water than alcohol.

(v) What do you mean by dynamic equilibrium?

Ans. The stage at which the number of molecules evaporating will be equal to the number of molecules coming back to liquid. This state is called dynamic equilibrium.

(vi) Why are the rates of diffusion in liquids slower than that of gases?

Ans. The attractive forces among the gas molecules are less than that of liquid hence the rate of evaporation in gases in greater than liquids.

(vii) Why does rate of diffusion increase with increase of temperature?

Ans. Diffusion increases by increasing temperature because at high temperature the intermolecular forces are weak and rate of diffusion increases.

(viii) Why are the liquids mobile?

Ans. The attractive forces among the liquid molecules are less than that of solid, hence they are mobile i.e. move freely.

5.4 TYPICAL PROPERTIES OF SOLID STATE

Q12. Write brief explanation of given properties of solids.

(i) Melting point (ii) Rigidity (iii) Density

Ans. (i) Melting point: The temperature at which the solid starts melting and coexists in dynamic equilibrium with liquid state is called melting point.

Explanation: The solid particles possess only vibrational kinetic energy. When solids are heated, their vibrational energies increase and particles vibrate at their mean position with a higher speed. If the heat is supplied continuously, a stage reaches at which the particles leave their fixed positions and then become mobile. At this temperature solid melts.

(ii) Rigidity: The particles of solids are not mobile. They have fixed positions. Therefore, solids are rigid in their structure.

(iii) Density of solids: Solids are denser than liquids and gases because solid particles are closely packed and do not have empty spaces between their particles. Therefore, they have the highest densities among the three state of matter. For example, density of aluminum is $2.70 gcm^{-3}$, iron is $7.86 gcm^{-3}$ and gold is $19.3 gcm^{-3}$.

5.5 TYPES OF SOLIDS

Q13. What are solids? Describe the classification of solids.

Ans. Solids: The state of matter which has definite shape and volume is called solid state e.g. book.

Classification of solids or types of solid: According to their general appearance solids

can be classified into two types. Amorphous solids and crystalline solid.

Amorphous solids: Amorphous means shapeless. Solids in which the particles are not regularly arranged or their regular shapes are destroyed, are called amorphous solids. They do not have sharp melting points. Plastic, rubber and even glass are amorphous solids as they do not have any sharp melting points.

Crystalline solids: Solids in which particles are arranged in definite three-dimensional pattern are called crystalline solids. They have definite surfaces or faces. Each face has definite angle with the other. They have sharp melting points. Example of crystalline solids are diamond, sodium chloride etc.

5.6 ALLOTROPY

Q14. (a) Define allotropy with examples.

(b) What is meant by transition temperature? Explain with examples.

Ans. (a) Allotropy: Two or more forms of same element having same chemical properties but different physical properties are called allotropes and the phenomenon is called allotropy.

OR

The existence of an element in more than one forms in same physical state is called allotropy.

Examples: Diamond and graphite are two allotropic forms of carbon.

Causes of allotropy:

(i) The existence of two or more kinds of molecules of an element each having different number of atoms e.g. oxygen (O_2) and ozone (O_3) are allotrope of oxygen.

(ii) Different arrangement of two or more atoms or molecules in a crystal of the element e.g. sulphur shows allotropy due to different arrangement of molecules (S_8) in the crystal.

Allotropes of solids have different arrangement of atoms in space at a given temperature. The arrangement of atoms can be changed with change in temperature and new allotropic form is produced.

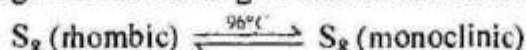
(b) Transition temperature:

The temperature at which one allotrope changes into another is called transition temperature.

Examples:

1. Transition temperature of sulphur is 96°C.

Below this temperature rhombic form is stable, if rhombic form is heated up to 96°C, its molecules arrange themselves to give monoclinic form.

$$S_8 \text{ (rhombic)} \xrightleftharpoons{96^\circ C} S_8 \text{ (monoclinic)}$$

2. Phosphorus $P_4 \text{ (white)} \xrightleftharpoons{250^\circ C} (P_4)_n \text{ (red)}$

3. Tin (grey) (cubic) $\xrightleftharpoons{18^\circ C}$ Tin (white) (tetrogonal)

White phosphorous:

White phosphorous is very reactive, poisonous and waxy soft solid. It exists as tetra-atomic molecules.

Red Phosphorous:

Red phosphorous is less reactive, non- poisonous and brittle powder.

Test yourself 5.5

(i) Which form of sulphur exists at room temperature?

Ans. Rhombic form of sulphur exists at room temperature.

(ii) Why is white tin available at room temperature?

Ans. White tin is available at room temperature because it exists at above 18°C and room temperature (25°C) is greater than its transition temperature. Hence it exists as white tin.

(iii) Why is the melting point of a solid considered its 'identification' characteristic?

Ans. Pure solids have fixed melting point which can not be changed, hence it can be considered its identification characteristic.

(iv) Why amorphous solids do not have sharp melting points while crystalline solids do have?

Ans. Amorphous solids do not have sharp melting points because their particles are not regularly arranged.

(v) Which is lighter one aluminium or gold?

Ans. Aluminium is lighter one.

(vi) Write the molecular formula of a sulphur molecule?

Ans. S_8 is a molecular formula of sulphur molecule.

(vii) Which allotropic form of carbon is stable at room temperature (25 °C)?

Ans. Diamond and graphite are stable at room temperature.

(viii) State whether allotropy is shown by elements or compounds or both?

Ans. Mostly the allotropy is shown by elements.

Science Technology Society:

Curing with salt to preserve meat: Table salt is the most important ingredient for curing meat and is used in large quantities. Salt kills and inhibits the growth of putrifying bacteria by drawing water out of the meat. Concentrations of salt up to 20% are required to kill most species of unwanted bacteria. Once properly salted, the meat contains enough salt to prevent the growth of many undesirable microbes.

Change of Instrumentation as the Science Progresses:

There are many aspects to be considered about the functioning of instruments. Scientific observation is determined by the human sensory system. It generally relies on instruments that serve as mediators between the world and the senses. Thus, instruments can be considered as a reinforcement of the senses. They provide a great capacity for increasing the power of observation and making induction processes easier. Furthermore, scientific instruments constitute a major factor in checking, refuting or changing previously established theories.

Key Points

- Gases diffuse very rapidly. Diffusion is mixing up of a gas throughout a space or other gases.
- Effusion is escaping of a gas molecule through a fine hole into an evacuated space.
- Gases exert pressure. The SI unit of pressure is Nm^{-2} which is also called Pascal.
- Standard atmospheric pressure is the pressure exerted by a mercury column of 760mm height at sea level, it is equivalent to 1 atmosphere.
- Gases are highly mobile and they can be compressed.
- Gases are 1000 times lighter than liquids or solids hence their density is measured in gdm^{-3}.
- Boyle's law states that volume of a given mass of a gas is inversely proportional to the pressure at constant temperature.
- Charles' Law states that volume of a given mass of a gas is directly proportional to the absolute temperature at a constant pressure.
- Absolute zero is the temperature at which an ideal gas would have zero volume, it is -273.15°C.
- The conversion of a liquid into vapours at all temperatures is called evaporation. It is a cooling process.
- Evaporation depends upon surface area, temperature and intermolecular forces.
- Vapour pressure of a liquid is defined as the pressure exerted by the vapours when liquid and vapour states are in dynamic equilibrium with each other.
- Boiling point is the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.
- Boiling point of a liquid is that temperature at which vapour pressure of liquid phase is equal to the vapour pressure of the solid phase. At this temperature liquid and solid coexist in dynamic equilibrium with one another.
- Melting point of solid is the temperature at which solid when heated melts and coexist in dynamic equilibrium with liquid.
- Solids are rigid and denser than liquids.
- Solids are classified as amorphous and crystalline.
- Amorphous solids are shapeless and do not have sharp melting point.
- Crystalline solids have definite three dimensional pattern of arrangement of particles. They have sharp melting points.
- The existence of a solid in different physical forms is called allotropy.

Exercise (Solved)

Multiple Choice Questions

Put a (✓) on the correct answer.

1. **How many times are the liquids denser than gases?**
(a) 100 times (b) 1000 times (c) 10,000 times (d) 100,000 times
2. **Gases are the lightest form of matter and their densities are expressed in terms of:**
(a) mg cm^{-3} (b) g cm^{-3} (c) g dm^{-3} (d) kg dm^{-3}
3. **Which one of the following coexists in dynamic equilibrium at freezing point:**
(a) gas and solid (b) liquid and gas (c) liquid and solid (d) all of these
4. **Which one of the following motions are possessed solid particles?**
(a) rotational motions (b) vibrational motions
(c) translational motions (d) both translational and vibrational motions
5. **Which one of the following is not amorphous?**
(a) rubber (b) plastic (c) glass (d) glucose
6. **One atmospheric pressure is equal to how many Pascals?**
(a) 101325 (b) 10325 (c) 106075 (d) 10523
7. **In the evaporation process, liquid molecules which leave the surface of the liquid have:**
(a) very low energy (b) moderate energy (c) very high energy(d) none of these
8. **Which one of the following gas diffuses fastest?**
(a) hydrogen (b) helium (c) fluorine (d) chlorine
9. **Which one of the following does not affect the boiling point?**
(a) intermolecular forces (b) external pressure
(c) nature of liquid (d) initial temperature of liquid
10. **Density of a gas increases, when its:**
(a) temperature is increased (b) pressure is increased
(c) volume is kept constant (d) none of these
11. **The vapour pressure of a liquid increases with the:**
(a) increase of pressure (b) increase of temperature
(c) increase of intermolecular forces (d) increase of polarity of molecules

Ans. 1. 1000 times	**2.** g dm^{-3}	**3.** liquid and solid	**4.** vibrational motions
5. glucose	**6.** 101325	**7.** very high energy	**8.** hydrogen
9. initial temperature of liquid	**10.** none of these	**11.** increase of temperature	

Short Answer Questions.

1. What is diffusion? Explain with an example.

Ans. For answer see Q.2

2. Define standard atmospheric pressure. What are its units? How it is related to Pascal?

Ans. For answer see Q.2

3. Why are the densities of gases lower than that of liquids?

Ans. For answer see Q.2

4. What do you mean by evaporation how it is affected by surface area?

Ans. **Evaporation:** The continuous escape of the molecules of a liquid from its surface is called evaporation and it occurs at any temperature.

Affect of surface area: Evaporation is a surface phenomenon. Greater is the surface, greater is evaporation and vice versa. e.g. sometimes a saucer is used if tea is to be cooled quickly.

5. Define the term allotropy with examples.

Ans. The existence of an element in more than one forms in same physical state is called allotropy.

Example 1.

Oxygen is present in two allotropic forms, O_2 and O_3.

Example 2.

Phosphorus is present in two allotropic forms. P_4 (white) and (P_4) red.

6. In which form sulphur exists at 100°C.

Ans. The surlphur exists in monoclinic form at 100 °C.

7. What is the relationship between evaporation and boiling point of a liquid?

Ans.

Evaporation	Boiling point
1. The continuous escape of the molecules of a liquid from its surface is called evaporation.	1. The change of liquid into the gaseous state at a particular temperature is called boiling.
2. It occurs at any temperature.	2. It occurs at fixed temperature.

Long Answer Questions

Q.1 Define Boyle's law and verify it with an example.

Ans. For answer see Q. 3.

Q.2 Define and explain Charles' law of gases.

Ans. For answer see Q. 4.

Q.3 What is vapour pressure and how it is affected by intermolecular forces?

Ans. For answer see Q. 7.

Q.4 Define boiling point and also explain, how it is affected by different factors.

Ans. For answer see Q.8.

Q.5 Describe the phenomenon of diffusion in liquids along with factors which influence it.

Ans. For answer see Q. 10.

Q.6 Differentiate between crystalline and amorphous solids.

Ans. For answer see Q. 13.

Numericals

Q.1 Convert the following units:

(a) 850mmHg to atm

Sol. 760 mm Hg = 1 atm

$$1 \text{ mm Hg} = \frac{1\text{atm}}{760}$$

$$850 \text{ mm} = \frac{1}{760} \times 850 \text{ atm}$$

$$= \frac{850}{760} \text{atm} = \mathbf{1.11842 \text{ atm.}}$$

(b)205000 Pa to atm

Sol. 101325 Pa = 1 atm

$$1\text{Pa} = \frac{1}{101325} \text{atm}$$

$$205000 \text{ Pa} = \frac{1}{101325} \times 205000 \text{atm}$$

$$= \frac{25000}{101325} \text{atm} = \mathbf{2.023 \text{ atm}}$$

(c) 560 torr to cm Hg

Sol. 760 torr = 76cmHg.

$$1 \text{ torr} = \frac{76}{760} \text{cmHg}$$

$$560 \text{ torr} = \frac{76}{760} \times 560 \text{cm Hg} = \mathbf{56 \text{ cm Hg}}$$

(d) 1.25 atm to Pa

Sol. 1 atm = 101325 Pa

1.25 atm = 101325×1.25 Pa

= **126656.25 Pa**

Q.2 Convert the following units:

(a) 750°C to K

Sol. As

K = °C + 273

∴ K = 750 + 273 = **1023 K**

(b) 150°C to K

Sol. As

K = °C+273

∴ K = 150+273 = **423 K**

(c) 100K to °C

Sol. As

°C = K–273 = 100 – 273 = **–173 °C**

(d) 172K to °C

Sol. As

°C = K–273 = 172 – 273 = **–101 °C**

Q.3 A gas at pressure 912mm of Hg has volume 450cm³. What will be its volume at 0.4atm.

Sol. Initial pressure $= P_1 = 912\text{mHg} = \frac{912}{760} = 1.2\text{atm}$

Final pressure $= P_2 = 0.4$ atm

Initial volume $= V_1 = 450\text{cm}^3$

Final volume $= V_2 = ?$

As $P_1V_1 = P_2V_2$

$\therefore \frac{P_1V_1}{P_2} = V_2$

OR $V_2 = \frac{P_1V_1}{P_2}$ Put the values

$V_2 = \frac{1.2\times450}{0.4} = 1350\text{cm}^3$

Result: The volume of gas = **1350cm³**

Q.4 A gas occupies a volume of 800cm³ at 1 atm, what will be its pressure in mm of Hg. When it is allowed to expand up to 1200cm³.

Sol. Initial volume $= V_1 = 800\text{cm}^3$

Final volume $= V_2 = 1200cm^3$

Initial pressure $= P_1 = 1$ atm $= 760$mm of Hg

Final pressure $= P_2 = ?$

As $P_1V_1 = P_2V_2$

$\therefore \frac{P_1V_1}{P_2} = P_2$

$P_2 = \frac{P_1V_1}{P_2}$ Put the values

$P_2 = \frac{760 \times 800}{1200}$

$= 506.66$

Result: Pressure of gas = **506.66mm of Hg**

Q.5 It is desired to increase the volume of a fixed amount of gas from 87.5 to 118cm³ while holding the pressure constant. What would be the final temperature if the initial temperature is 23°C?

Sol. Initial volume $= V_1 = 87.5cm^3$

Final volume $= V_2 = 118cm^3$

Initial temperature $= T_1 = 23°C = 23-273 = 296K$

Final temperature $= T_2 = ?$

As $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

OR $V_1T_2 = V_2 = T_1$

OR $T_2 = \frac{V_2T_1}{V_1}$ Put the values

$T_2 = \frac{118 \times 296}{87.5} = 399.177$ K

Result: Final temperature = **399.177 K**

$K = C° + 273$

$K - 273 = °C$

$C° = K - 273$

$C° = 399.177 - 273 = 126$

Q.6 A sample of gas is cooled at constant pressure from 30°C to 10°C. Comment:

a. Will the volume of the gas decrease to one third of its original volume?

b. If not, then by what ratio will the volume decrease?

Sol.(a) Yes, the volume will decrease.

(b) Yes, volume will decrease to one third of its original volume (according to Charle's law)

Q.7 A balloon that contains 1.6 dm³ of air at standard temperature (0°C) and (1atm) pressure is taken under water to a depth at which its pressure increases to 3.0 atm. Suppose that temperature remain unchanged, what would be the new volume of the balloon. Does it contract or expand?

Sol. Initial volume $= V_1 = 1.6\ dm^3$

Final volume $= V_2 = ?$

Initial pressure $= P_1 = 1\ atm$

(Standard pressure)

Final pressure $= P_2 = 3.0\ atm$

As $P_1V_1 = P_2V_2$

$\therefore \frac{P_1V_1}{P_2} = V_2$

OR $V_2 = \frac{P_1V_1}{P_2}$ Put the values

$V_2 = \frac{1\times1.6}{3} = 0.533\ dm^3$

Result: Volume of balloon = **0.533 dm³**

Hence it will contract.

Q.8 A sample of neon gas occupies a volume 75.0cm³ at very low pressure of 0.4 atm. Assuming temperature remains constant what would be the volume at 1.0 atm. pressure?

Sol. Initial volume $= V_1 = 75.0\ cm^3$

Final volume $= V_2 = ?$

Initial pressure $= P_1 = 0.4\ atm$

Final pressure $= P_2 = 1\ atm$

As $P_1V_1 = P_2V_2$

$\therefore \frac{P_1V_1}{P_2} = V_2$

$V_2 = \frac{P_1V_1}{P_2}$ Put the values

$V_2 = \frac{0.4\times75}{1} = 30\ cm^3$

Result: Volume of gas = **30 cm³**

Q.9 A gas occupies a volume of 35.0 dm³ at 17°C. If the gas temperature rises to 34°C at constant pressure, would you expect the volume to double? If not calculate the new volume.

Sol. Initial volume $= V_1 = 35.0\ dm^3$

Final volume $= V_2 = ?$

Initial temperature $= T_1 = 17°C = 17+273 = 290k$

Final temperature $= T_2 = 34°C = 34+273 = 307k.$

As $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$\therefore \frac{V_1}{T_1} \times T_2 = V_2$

OR $V_2 = \frac{V_1 \times T_2}{T_1}$ Put the values

$V_2 = \frac{35 \times 307}{290}$

$= 37.05\ dm^3$

Result: Volume of gas = **37.05 dm³**

Q.10 The largest moon of Saturn, is Titan. It has atmospheric pressure of 1.6×10⁵ Pa. What is the atmospheric pressure in atm? Is it higher than earth's atmospheric pressure?

Sol. Pressure $= 1.6\times10^5$ Pa

101325 Pa = 1 atm

1 Pa $= \frac{1}{101325}$ atm

1.6×10^5 Pa $= \frac{1}{101325} \times 1.6\times10^5$

$= \frac{1.6\times10^5}{101325}$

$= 1.579$ atm

Result: Atmospheric pressure in atm = **1.579**

Yes, it is higher than atmosphere present of earth

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

5.1	Typical Properties of Gaseous State
5.2	Laws Related to Gases
5.3	Typical Properties of Liquid State

☆ Tick the correct answer.

1. **The simplest form of matter is:** (LHR. GI, MLN. GII, MLN. GI)
   (A) Gas (B) Liquid (C) Solid (D) Both B and C
2. **One atmospheric pressure is equal to how many Pascal's:** (LHR. GII)
   (A) 10325 (B) 101325 (C) 106075 (D) 10523
3. **The density of ice at 0°C is:** (FBD. GI)
   (A) $1.00gcm^{-3}$ (B) $1.5gdm^{-3}$ (C) $0.917gcm^{-3}$ (D) $1.4gdm^{-3}$
4. **Density of gases is expressed in terms of:** (SWL. GI, SGD. GII, MLN. GII)
   (A) $mgcm^{-3}$ (B) gcm^{-3} (C) gdm^{-3} (D) $kgdm^{-3}$
5. **SI unit of pressure is:** (SWL. GII)
   (A) Nm^{-2} (B) $N^{-2}m$ (C) $N^{-1}m^{2}$ (D) Nm
6. **Which one of the following gas diffuses fastest?** (SGD. GI, BWP. GI)
   (A) Hydrogen (B) Helium (C) Flourine (D) Chlorine
7. **Density of a gas increases when:** (RWP. GII)
   (A) Temperature increases (B) Pressure increases
   (C) Volume is kept constant (D) Volume increases
8. **Tyre puncture is an example of:** (LHR. GII)
   (A) Effusion process (B) Diffusion process
   (C) Evaporation process (D) Condensation process
9. **Density of aluminium is:** (GRW. GI)
   (A) 2.4 g cm^{-3} (B) 2.5 g cm^{-3} (C) 2.6 g cm^{-3} (D) 2.7 g cm^{-3}
10. **The value of atmospheric pressure at sea level is:** (GRW. GII)
    (A) 760 mm Hg (B) 700 mm Hg (C) 780 mm Hg (D) 750 mm Hg
11. **The apparatus used to measure atmospheric pressure is:** (SWL. GII)
    (A) thermometer (B) galvanometer (C) ammeter (D) barometer
12. **Gases can be compressed because:** (BWP. GI)
    (A) there are no empty spaces between gas molecules

(B) there are large empty spaces between gas molecules
(C) molecules are very close to each other
(D) molecules are very large in size

13. In Charles Law "K" is equal to: (LHR. GI, BWP. GII)
(A) $\frac{T}{V}$ (B) TV (C) $\frac{V}{T}$ (D) $\frac{V}{P}$

14. The constant factor in Bolyle's law is: (SWL. GI)
(A) volume (B) pressure (C) temperature (D) mole

15. How many times liquids are denser than gases? (DGK. GI)
(A) 100 times (B) 1000 times (C) 10000 times (D) 100000 times

16. The blood pressure of a healthy man is: (GRW. GI, DGK. GI)
(A) $\frac{120}{80}$ mm Hg (B) $\frac{140}{90}$ mm Hg (C) $\frac{110}{100}$ mm Hg (D) $\frac{150}{70}$ mm Hg

17. Normal body temperature of human being is: (GRW. GII)
(A) 37°C (B) 38°C (C) 39°C (D) 40°C

18. The freezing point of water is: (FBD. GII)
(A) 2°C (B) 0°C (C) 1°C (D) 100°C

19. The Vapour Pressure of a liquid increases with increase of: (MLN. GI, DGK. GII)
(A) Pressure (B) Temperature
(C) Intermolecular Forces (D) Polarity of Molecules

20. How many times liquids are denser than gases? (RWP. GII, DGK. GI)
(A) 1000 times (B) 100 times (C) 10,000 times (D) 100,000 times

21. Freezing point of acetic acid is: (FBD. GI)
(A) 14.6°C (B) 15.6°C (C) 16.6°C (D) 17.6°C

22. The boiling point of water is: (FBD. GII)
(A) 0°C (B) 60°C (C) 100°C (D) 120°C

23. By increasing temperature the rate of evaporation is: (RWP. GII)
(A) increased (B) decreased (C) becomes equal (D) not effected

Answers

1. Gas
2. 101325
3. $0.917 gcm^{-3}$
4. gdm^{-3}
5. Nm^{-2}
6. Hydrogen
7. Pressure increases
8. Effusion process
9. $2.7 g cm^{-3}$
10. 760 mm Hg
11. barometer
12. there are large empty spaces between gas molecules
13. $\frac{V}{T}$
14. temperature
15. 1000 times
16. $\frac{120}{80}$ mm Hg
17. 37°C
18. 0°C
19. Temperature
20. 1000 times
21. 16.6°C
22. 100°C
23. increased

☆ Give short answer to the following questions.

1. Define pressure. Write down its unit. (LHR. GI, BWP. GII, RWP. GI, MLN. GI)

Ans. Pressure: The force exerted per unit surface area (A) is called pressure.

That force which is exerted by a gas on per unit area is called its pressure. The pressure is represented by (P).

$P = F/A$

The S.I unit of force is Newton, and unit of area is m^2, so the unit of pressure is Nm^{-2}, it is also known as Pascal. $P = 1Nm^{-2}$

2. Define standard atmospheric pressure. (GRW. GII)

Ans. A pressure exerted by atmosphere at the sea level. It can be defined as. The pressure exerted by a mercury column of 760mm height at sea level."

1 atm = 760mm of Hg

= 760 torr (1mm of Hg = One torr)

= 101325 Nm^{-2}

= 101325 Pa

3. Define the term effusion and give one example. (FBD. GI, FBD. GII)

Ans. The escaping of gas molecules through a tiny hole into a space toward area of low pressure. e.g when a tyre get punctured, air effuse out.

4. Differentiate between diffusion and effusion. (SWL. GII, FBD. GI, MLN. GI)

Ans. Diffusion: "Spontaneous mixing up of gas molecules by random motion and colluision is called diffusion". Diffusion depends upon the molecular mass of gases. Lighter gases diffuse rapidly.

Effusion: "Escape of gas molecule through a tiny hole to an evacuated space".

e.g. when a tyre get punctured, all the air effuses out. Effusion also depend upon molecular mass.

5. Why the gases are compressible? (SGD. GI, DGK. GI)

Ans. The molecules of gases are farther apart. There is blank spaces among them, so by compressing, these spaces reduced.

6. Why the density of gases is lesser than that of liquids? (SGD. GII)

Ans. Because the gas molecules have very large spaces among them. Hence their light mass and more volume; gases have low densities.

7. Whether the density of gases increases on cooling? (RWP. GII)

Ans. On cooling the molecules of gases come close to each other, so the density is increased by reducing the volume.

8. Which factors affect diffusion of liquids? (GRW. GII)

Ans. The diffusion of liquid depend upon following factors.

1. **Intermolecular forces:** The rate of diffusion is high in those molecules

who have weak intermolecular forces.

2. **Size of molecules:** Large size molecules have slow rate of diffusion.
3. **Shape of molecules:** The molecules of proper shapes can be easily diffused.
4. **Temperature:** Higher the temperature, higher the diffusion will be.

9. What is meant by freezing point? (MLN. GII)

Ans. That temperature at which vapour pressure of liquid phase is equal to vapour pressure of solid phase. At this temperature liquid and solid coexist in dynamic equilibrium with one another.

10. Why is the density of gas measured in gdm^{-3} while that of liquid in gcm^{-3}? (RWP. GI)

Ans. The density of gas is less than that of liquid, that is why density of gas is expressed in gdm^{-3} and that of liquid in gcm^{-3}.

11. On which two factors evaporation depends on? (RWP. GII)

Ans. Evaporation depends upon following factors

1. Surface area 2. temperature

12. What do you mean by pascal? How many pascals is equal to 1 atm? (BWP. GII)

Ans. If one Newton force exerted perpendicular to area of 1 meter square, then the pressure exerted on a body is equals to (1)one pascal.

1 atm = 101325 Pa

13. Define Charle's law. (LHR. GII. RWP. GI. MLN. GI)

Ans. Charles law: The law was put forward by a French Scientist J. Charles in 1787.

This law states that:

The volume of given mass of a gas is directly proportional to the absolute temperature at constant pressure.

$$V \propto T$$

$$V = kT; \frac{V}{T} = k$$

14. Differentiate between Boyle's law and Charles' law. (GRW. GII)

Ans.

	Boyle's law	Charle's law
(i)	Boyles studied the relation between gas volume and pressure at constant temperature.	Charles studied the relation between gas volume and temperature at constant pressure.
(ii)	He observed "If temperature is kept constant, the volume of mass of gas is inversaly proportional to its pressure."	He observed "If pressure is kept constant, the volume of mass of a gas is directly proportional to it temperature."

15. Define Boyle's law and write down its mathematical expression.
(SWL. GI, FBD. GII, MLN. GI, BWP. GI, GRW. GI, DGK. GI, SGD. GII)

Ans. In 1662, Robert Boyle studied the relationship between volume and pressure at

constant temperature.

According to law:

"If temperature is kept constant, the given mass of gas is inversaly proportional to its pressure".

$V \propto \frac{1}{P}$

or $V = \frac{k}{P}$ or $VP = k$

16. What is absolute zero temperature? (LHR. GII, SWL. GI)

Ans. That temperature at which the volume of an ideal gas become zero. Its value is -273.15°.

17. Convert 50°C into Kelvin scale. (DGK. GII)

Ans.

$$K = T°C + 273$$
$$= 50°C + 273$$
$$= 323k$$

18. Why vapour pressure is higher at high temperature? (LHR. GI, RWP. GII)

Ans. At high temperature, vapour pressure is higher than at low temperature. Because at elevated temperature, the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

19. What is meant by evaporation? (LHR. GII, RWP. GII, FBD. GII, BWP. GI)

Ans. The process in which liquid state changes in vapour state is called evaporation. This process occurs only at liquid surface. Greater the surface of liquid greater will be evaporation.

20. What is meant by condensation? (GRW. GI, MLN. GII)

Ans. Condensation: A process in which gaseous phase changes into liquid phase is called condensation.

21. Why are the rates of diffusion in liquids slower than that of gasses? (GRW. GII, MLN. GI)

Ans. In gases, the molecules moves freely and have less force of attraction among them. So they occupy all the available space. Instead in liquids molecules have strong force of attraction as compared to gases so they are less mobile, that's the reason why liquids diffuse slowly.

22. Why does evaporation increase with increase of temperature? (FBD. GII)

Ans. At high temperature, rate of evaporation becomes fast, because at high temperature, the kinetic energy of molecules become very high and over come the intermolecular forces. For example water level in hot water container decreases earlier than that of cold water. This is because the hot water evaporate earlier.

23. Why are the liquids are mobile? (SWL. GII, DGK GII)

Ans. Among the molecules of liquid, intermolecular forces are not much strong, to keep them stationary, instead the molecules keeps on moving freely under the surface of liquid. That's why liquid have not definite shape but have definite volume.

24. Why is the boiling point of water higher than that of alcohol? (SGD. GI)

Ans. The intermolecular forces in water are much strong than that of alcohol, that is the reason why water has high boiling point.

25. Why evaporation causes cooling? (RWP. GI, DGK. GII, LHR. GI, GRW. GII)

Ans. Evaporation is a cooling process, when the high kinetic energy molecules vapourize, the temperature of remaining molecule fall down. To compensate this deficiency of energy, molecule of liquid absorb energy from surrounding. As a result temperature of surrounding decreases, and we cool effect.

26. Write down two factors on which Vapour Pressure depends upon. (BWP. GI)

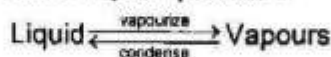
Ans. The vapour pressure of liquid depend upon following factor.

Molecular size: The small size molecules vapourize rapidly as compared to big ones.

Temperature: As compare to low temperature, at high temperature vapour pressure is high.

27. Define the term vapour pressure. (LHR. GI, SGD. GII)

Ans. The pressure exerted by the vapours of a liquid at equilibrium with the liquid at particular temperature is called vapour pressure.

$$\text{Liquid} \underset{\text{condense}}{\overset{\text{vapourize}}{\rightleftharpoons}} \text{Vapours}$$

28. Convert 70cm Hg to atm. (FBD. GI)

Ans. By converting 70cm Hg to atm

76cm Hg = 1atm

1cm Hg = $\frac{1}{76}$ atm

70cm Hg = $\frac{1}{76} \times 70$ = 0.92atm

29. Define boiling point. What is boiling point of alcohol? (LHR. GI, FBD. GI)

Ans. The temperature at which vapour pressure of liquid become equal to external pressure.

Boiling point of alcohol is 78°C.

30. Why meat is preserved curing with salt? (FBD. GII, SGD. GI)

Ans. Table salt is the most important ingredient for curing meat and is used in large quantity. Salt kills and inhibit the growth of putrefying bacteria by drawing water out of meat. Concentration of salt upto 20% is required to kill most species of

unwanted bacteria. Once properly salted, the meat contains enough salt to prevent the growth of many undesirable microbes.

31. What is relationship between external pressure and boiling point? (SWL. GI)

Ans. The boiling point of liquid is directly proportional to external pressure. With the increase of external pressure, boiling point also increased.

32. Name the factors which affect the vapour pressure of liquid. (SWL. GI)

Ans. Following are the factors

i. Nature of liquid ii. Molecular size iii. Temperature

33. Convert the 3.5 atm to torr. (SGD. GI)

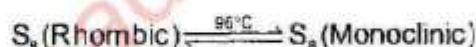
Ans. 3.5 atm to torr

1 atm = 760 torr

3.5 atm = 3.5 × 760 = 2660 torr

34. Define transition temperature and give an example. (SGD. GI)

Ans. A temperature at which one allotrope is converted to another is known as transition temperature. For example transition temperature of sulphur is 96°C. Below this temperature it is found in rhombic form. If rhombic form again heated till 96°, it is change into monoclinic form.

$$S_8(\text{Rhombic}) \xrightleftharpoons{96^\circ C} S_8(\text{Monoclinic})$$

35. Convert 700 mm of Hg into atmosphere (atm). (DGK. GII)

Ans. 760mm Hg = 1atm

$$1\text{mm Hg} = \frac{1\text{atm}}{760}$$

$$700\text{mm} = \frac{1\text{atm}}{760} \times 700\text{mmHg} = 0.92\text{atm}$$

36. Define Systolic pressure. (BWP. GII)

Ans. When heart is pumping, the value of blood pressure that represent this value of pressure is known as systolic pressure. e.g 120

5.4	Typical Properties of Solid State
5.5	Types of Solids
5.6	Allotropy

☆ Tick the correct answer.

1. Solid particles possess which one of the following motion? (RWP. GI)

(A) Rotational motion (B) Vibrational motion

(C) Translational motion (D) Simple motion

2. **The temperature at which an ideal gas would have zero volume is:** (SGD. GI)
(A) –760°C (B) –173.15°C (C) –273.15°C (D) 0°C

3. **Freezing point of ethyl alchol is:** (RWP. GI)
(A) +115°C (B) –115°C (C) –116°C (D) +116°C

4. **Which is not amorphous?** (FBD. GI, BWP. GI, RWP. GI, DGK. GII)
(A) Rubber (B) Plastic (C) Glass (D) Glucose

5. **Which one is not solid amorphous?** (FBD. GII)
(A) Rubber (B) Plastic (C) Glass (D)Sodium chloride

6. **How much concentrations of salt is required to kill unwanted bacteria?** (SGD. GI)
(A) 5% (B) 10% (C) 15% (D) 20%

Answers

1.	Vibrational motion	2.	–273.15°C	3.	–115°C
4.	Glucose	5.	Sodium chloride	6.	20%

☆ **Give short answer to the following questions.**

1. **Define the term melting point.** (FBD. GI)

Ans. The temperature at which solid start melting and coexist in dynamic equilibrium with liquid state is called melting point.

2. **Differentiate between boiling point and melting point.** (SWL. GI)

Ans.

Melting point	Boiling point
A temperature at which solid starts melting, and coexist with liquid state at dynamic equilibrium.	A temperature at which the temperature of liquid becomes equal to atmospheric pressure.

3. **What is the effect of temperature on density of gases?** (GRW. GII)

Ans. By lowering the temperature, the volume of gases decreases, while density increased. In normal atmospheric pressure, the density of oxygen at 20°C is $1.4 gdm^{-3}$ and at 0°C it is $1.5 gdm^{-3}$.

4. **Differentiate between amorphous solids and crystalline solids.** (LHR. GI, RWP. GI)

Ans. Amorphous Solids: Amorphous means shapeless. Solids in which the particles are not regularly arranged or those having not regular shapes are called amorphous solids. For example: Rubber, Plastic.

Crystalline Solids:

Solids in which particles are arranged in definite three-dimensional pattern are called crystalline solids. For example: salt and diamond.

5. What is meant by Amorphous Solid? (MLN. GI, DGK. GII, GRW. GI)

Ans. Amorphous means shapeless. The type of solids in which particles are not properly arranged or those having not proper shapes are called amorphous solids.
e.g. Plastic, rubber

6. Define crystalline solid and give its two example. (DGK. GI)

Ans. Those solids whose particles are organized three-dimensionally called crystalline solids. They have high melting and boiling point. e.g. Diamond, Salt.

7. Define Allotropy. (BWP. GI, LHR. GII, DGK. GI)

Ans. The existance of an element in two or more form in same physical state is known as allotropy.

8. Write two reasons of allotropy. (GRW. GI)

Ans. Two reasons of allotropy are following:

(1) Occurance of element in two or more forms, having different number of atoms i.e allotrop of oxygen O_2 and O_3.

(2) Different arrangement of molecules or atoms in crystal of element like sulphur crystal show allotropy due to different arrangement of molecules.

9. Why solids show rigidity? (GRW. GI)

Ans. Due to strong intermolecular forces the molecules of solid do not move, they are tightly packed. That is why solids have specific shape.

10. What are the physical properties of matter? (MLN. GII)

Ans. The simplest form of matter is gas.

1. In gaseous state, matter has no specific shape and volume.
2. Gas molecules are mobile, they can easily pressed.
3. In liquid form, matter has specific volume, but not shape.
4. By heating, their volume is increased.
5. In solid form matter has specific shape and volume.
6. On heating solids melts and change into liquid or gaseous state.

11. Define diffusion, explain with an example. (LHR. GII, BWP. GII, SGD. GII)

Ans. Spontaneous mixing up of gas molecules, by random motion and collosion to make homogeneous mixture, called diffusion.

Diffusion depend upon the molecular mass of gases, light gases diffuse rapidly e.g hydrogen (H_2) gas diffuses four times rapidly than oxygen (O_2) gas.

Major Concepts:

6.1	Solution, aqueous solution, solute and solvent	
6.2	Saturated, unsaturated, supersaturated solutions and dilution of solution	
6.3	Types of solutions	6.4 Concentration units
6.5	Comparison of solutions, suspensions and colloids	

Time allocation	
Teaching periods	16
Assessment periods	02
Weightage	14%

Students Learning Outcomes:

Students will be able to:

- Define the terms: solution, aqueous solution, solute and solvent and give an example of each.
- Explain the difference between saturated, unsaturated and supersaturated solutions.
- Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each.
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each.
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each.
- Explain what is meant by the concentration of a solution.
- Define molarity.
- Define percentage solution.
- Solve problems involving the molarity of solution.
- Describe how to prepare dilute solutions from concentrated solutions of known molarity.
- Convert between the molarity of a solution and its concentration in g/dm^3.
- Use the rule that "like dissolves like" to predict the solubility of one substance in another.

6.1 SOLUTION
6.2 SATURATED SOLUTION

Q.1 Define the following.

(i) Solution (ii) Solute (iii) Solvent (iv) Binary solution
(v) Aqueous solution (vi) Saturated solution

(vii) Super saturated solution *(viii) Unsaturated solution*
(ix) Dilute solution *(x) Concentrated solution*

Ans. (i) Solution: A homogeneous mixture of different chemical substances which has uniform chemical composition through out and shows uniform physical properties is called solution. **OR**

Homogeneous mixture of two or more than two substances is called solution e.g. sugar solution (sugar + water)

(ii) Solute: The component of solution present in relatively lesser amount is called solute e.g. in sugar solution (sugar + water), sugar is solute.

(iii) Solvent: The component of solution present in relatively large amount in solution is called solvent e.g. in sugar solution [sugar + water], water is a solvent.

(iv) Binary solution: A solution which is formed by mixing only two substances is called binary solution e.g. sugar solution [sugar + water]

(v) Aqueous solution: The solution which is formed by dissolving substance in water is called an aqueous substance. **OR**

The solution whose solvent is water is called aqueous solution.

Examples: Solution of sugar in water is aqueous solution of sugar and solution of salt in water is aqueous solution of salt.

(vi) Saturated solution: A solution containing maximum amount of solute at a given temperature is called saturated solution. In saturated solution undissolved solute is in equilibrium with dissolved solute.

$$\text{Solute (crystallized)} \rightleftharpoons \text{solute (dissolved)}$$ **OR**

The solution which can not dissolve more of the solute at a particular temperature is called a saturated solution at that temperature.

(vii) Supersaturated solution: A solution which contains more amount of the solute than that required for preparing its saturated solution at a particular temperature. **OR**

The solution which is more concentrated than a saturated solution is called supersaturated solution.

Explanation: When saturated solutions are heated they develop further capacity to dissolve more solute. Such solutions contain greater amount of solute than, required to form a saturated solution and they became more concentrated.

(viii) Unsaturated solution: A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution. Such solutions have the capacity to dissolve more solute to become a saturated solution.

OR

A solution which can dissolve further amount of a solute at a particular temperature is called an unsaturated solution.

(ix) Dilute Solution: A solution containing low concentration of a solute is called a

dilute solution.

(x) Concentrated solution: A solution containing high concentration of the solute is called concentrated solution.

6.3 TYPES OF SOLUTION

Q.2 Describe the various types of solutions with examples.

Ans. Types of solution: Each solution consists of two components solute and solvent. The solute as well as solvent may exist as gas, liquid or solid. So, depending upon the nature of solute and solvent different types of solutions may form which are given in below.

Different Types of solutions with examples

Sr. No.	Solute	Solvent	Example of Solution
1.	Gas	Gas	Air, mixture of H_2 and He in balloons, mixture of N_2 and O_2 in cylinders for respiration.
2.	Gas	Liquid	Oxygen in water, carbon dioxide in water.
3.	Gas	Solid	Hydrogen adsorbed on palladium.
4.	Liquid	Gas	Mist, fog, liquid air pollutants.
5.	Liquid	Liquid	Alcohol in water, benzene in toluene.
6.	Liquid	Solid	Butter, cheese.
7.	Solid	Gas	Dust particles or smoke in air.
8.	Solid	Liquid	Sugar in water.
9.	Solid	Solid	Metal alloys (brass, bronze), opals.

Test yourself 6.1:

i. Why is a solution considered mixture?

Ans. Solutions are homogeneous mixtures of two or more substances which can be separated by physical means and have sharp melting and boiling point.

ii. Distinguish between the following pairs as compound or solution:
(a) water and salt solution (b) vinegar and benzene (c) carbonated drinks and acetone.

Ans. (a) Water is a compound and salt solution is a mixture.
(b) Vinegar is a solution and benzene is a compound.
(c) Carbonated drink is a solution and acetone is a compound.

iii. What is the major difference between a solution and a mixture?

Ans.

Solution	Mixture
Solution is always homogeneous mixture	Mixture may be homogeneous or heterogeneous

iv. Why are the alloys considered solutions?

Ans. Alloys are considered solution because they are homogeneous, mixture of metals.

v. Dead sea is so rich with salts that it forms crystals when temperature lowers in the winter. Can you comment why is it named as "Dead Sea"?

Ans. It is called Dead sea because nothing can sink in it due to high density of salted water and life can out survive in it.

6.4 CONCENTRATION UNITS

Q.3(a) What is meant by concentration?

(b) What is percentage? Describe the various ways of expressing concentration of a solution.

Ans. Concentration: The amount of a solute which has been dissolved in a particular amount of a solvent is called concentration of solution. **OR**

Concentration is the proportion of a solute in a solution. It is also a ratio of amount of solute to the amount of solution or ratio of amount of solute to amount of solvent.

Note: Please keep in mind that concentration does not depend upon the total volume or total amount of the solution. For example a sample taken from the bulk solution will have the same concentration.

(b) Percentage: Percentage unit of concentration refers to the percentage of solute present in a solution. The concentration of a solution can be expressed in following ways.

1. Percentage mass/ mass (%m/m): It is the number of grams of solute in 100 grams of solution. For example, 10% m/m sugar solution means that 10g of sugar is dissolved in 90g of water to make 100g of solution.

$$\%\ \text{mass/mass} = \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100$$

$$= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

2. Percentage mass/ volume (%m/v): It is the number of grams of solute dissolved in $100cm^3$ (parts by volume) of solution. For example 10% m/v sugar solution contains 10g of sugar $100cm^3$ of solution. The exact volume of solvent is not mentioned or it is not known.

$$\%\text{m/v} = \frac{\text{Mass of solute(g)}}{\text{volume of solution } (cm^3)} \times 100$$

Percentage - volume/ mass (%v/m): It is the volume in cm^3 of a solute dissolved in 100g of the solution. For example 10% v/m alcohol solution in water means $10cm^3$ of alcohol is dissolved in (unknown) volume of water so that the total weight of solution is 100g. In such solutions the mass of solution is under consideration, total volume of the solution is not considered.

$$\%\ \text{v/m} = \frac{\text{volume of solute } (cm^3)}{\text{Mass of solution (g)}} \times 100$$

3. Percentage volume/ volume (% v/v): It is the volume in cm^3 of a solute dissolved per $100cm^3$ of the solution, For example, 30 percent alcohol solution means $30cm^3$ of alcohol dissolved in sufficient amount of water, so that the total volume of the solution becomes $100cm^3$.

$$\% \text{ volume/volume} = \frac{\text{volume of solute } (cm^3)}{\text{volume of solution } (cm^3)} \times 100$$

Example 6.1:

If we add 5cm³ of acetone in water to prepare 90cm³ of aqueous solution. Calculate the concentration (v/v) of this solution.

Solution: Using the relationship

$$\% \text{ volume/volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

$$= \frac{5}{90} \times 100 = 5.5$$

Thus concentration of solution is 5.5 percent by volume

Q.4.(a) Define the following.

(i) Molarity (ii) Molar solution

(b) Differentiate pure liquid and solution.

Ans.(a) (i) Molarity: Molarity is a concentration unit. it is defined as the

Number of moles of solute present in one dm^3 of the solution. It is represented by M.

$$\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{volume of solution in } dm^3}$$

As.

$$\text{Number of moles} = \frac{\text{mass of solute in grams}}{\text{molar mass of solute}}$$

$$M = \frac{\text{Mass of solute in grams}}{(\text{Molar mass of solute } (g\,mol^{-1})) \times \text{volume of solution in } (dm^3)}$$

(ii) Molar solution: A solution in which one mole of solute has been dissolved in one dm^3 of solution is called molar solution. It is represented by 1M.

Example: When one mole of sodium hydroxide (40g) is dissolved in one dm^3 of water, it will be the 1M solution of sodium hydroxide.

(b) Difference between pure liquid and a solution.

The simplest way to distinguish between a solution and a pure liquid is evaporation. The liquid which evaporates completely, leaving no residue, is a pure liquid while a liquid which leaves, behind a residue on evaporation is a solution.

Test yourself 6.2

i. Does the percentage calculations require the chemical formula of the solute?

Ans. No, only amounts of solute and solvent are required.

ii. Why is the formula of solute necessary for calculation of the molarity of the solution?

Ans. In order to determine the molarity molar mass in required, for which formula of solute is necessary.

iii. You are asked to prepare 15 percent (m/m) solution of common salt. How much amount of water will be required to prepare this solution?

Ans. 85g water will be required to prepare this solution.

iv. How much water should be mixed with $18cm^3$ of alcohol so as to obtain 18% (v/v) alcohol solution?

Ans. $82cm^3$ water should be mixed to obtain 18% v/v alcohol solution.

v. Calculate the concentration % (m/m) of a solution which contains 2.5g of salt dissolved in 50g of water.

Ans. $\%age\ (w/w) = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100 = 4.76$

$$= \frac{2.5}{2.5 + 50} \times 100$$

$$= \frac{2.5}{52.5} \times 100$$

vi. Which one of the following solutions is more concentrated:
One molar or three molar

Ans. Three molar solution is more concentrated.

Example 6.2:

Calculate the molarity of a solution which is prepared by dissolving 28.4g of Na_2SO_4 in $400cm^3$ of solution.

Solution:

Conversion mass of solute into moles

$$\text{No. of moles } Na_2SO_4 = \frac{\text{mass dissolved(g)}}{\text{molar mass } (gmol^{-1})}$$

$$= \frac{28.4g}{142\ gmol^{-1}}$$

$$= 0.2 \text{ mol}$$

$$\text{Conversion of volume into } dm^3 = \frac{400cm^3}{1000cm^3} \times 1dm^3$$

$$= \mathbf{0.4dm^3}$$

$$\text{Molarity} = \frac{\text{no. of moles}}{\text{volume of solution } (dm^3)}$$

$$= \frac{0.2}{0.4} = \mathbf{0.5\ mol\ dm^{-3}}$$

Example 6.3: How much NaOH is required to prepare its $500\ cm^3$ of 0.4M solution.

Solution: Molar mass of NaOH = 40 $gmol^{-1}$

$$\text{Volume in } dm^3 = \frac{500cm^3}{1000cm^3} \times 1dm^3 = 0.5dm^3$$

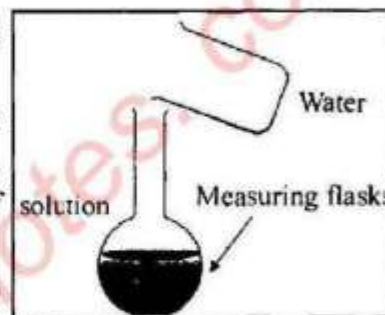
Putting the values in formula:

$$\text{Molarity} = \frac{\text{Mass of solute (g)}}{\text{molar mass (gmol}^{-1}) \times \text{volume of solution (dm}^3)}$$

Mass of solute = Molarity×molar mass×volume
= 0.4×40×0.5 = **8g**

Q.5. How can you prepare dilute solution (0.01M) from concentrated (0.1M) solution?

Ans. Dilute molar solution is prepared from a concentrated solution of known molarity as explained below.

Suppose we are to make 100 cm³ of 0.01M solution from given 0.1M solution of potassium permanganate. First 0.1 M solution is prepared by dissolving 15.8g of potassium permanganate in 1 dm³ of solution. Then 0.01 M solution is prepared by the dilution according to following calculations.

Dilution of a solution

Concentrated solution Dilute solution

$$M_1V_1 = M_2V_2$$

Where $M_1 = 0.1$ M
$V_1 = ?$

and

$V_2 = 100\text{cm}^3$
$M_2 = 0.01\text{M}$

Putting the values in above equation we get

Concentrated solution Dilute solution

$$V_1 \times 0.1 = 0.01 \times 100$$

$$V_1 = \frac{0.01 \times 100}{0.1}$$

$$= 10\text{cm}^3$$

Concentrated solution of $KMnO_4$ has dense purple colour. We take 10cm³ of this solution with the help of a graduated pipette and put in a measuring flask of 100cm³. Add water upto the mark present at the neck of the flask. Now it is 0.01 molar solution of $KMnO_4$.

Example 6.4:

10cm³ of 0.01 molar $KMnO_4$ solution has been diluted to 100cm³. Find out the molarity of this solution.

Solution:

Data: $M_1 = 0.01\text{M}$ $M_2 = ?$
$V_1 = 10\text{cm}^3$ $V_2 = 100\text{cm}^3$

Using following formula, volume required can be calculated

$$M_1V_1 = M_2V_2$$

or $$M_2 = \frac{M_1V_1}{V_2}$$

By putting these values we get molarity

$$M_2 = \frac{0.01 \times 10}{100} = 0.001M$$

6.5 SOLUBILITY

Q.6.(a) What is meant by solubility? Discuss the general principle of solubility.

(b) Explain the solute - solvent interaction for the preparation of solution.

Ans. Solubility: Solubility is defined as the number of grams of the solute dissolved in 100g of solvent to prepare a saturated solution at a particular temperature.

General Principle of Solubility: The general principle of solubility is like dissolves like.

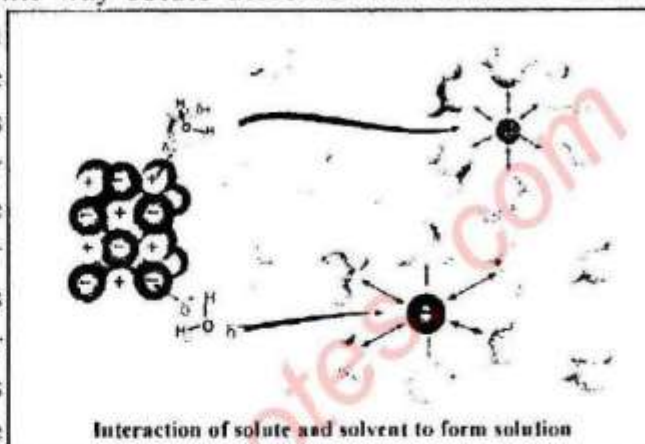
1. The polar substances are soluble in polar solvents. Ionic solids and polar covalent compounds are soluble in water e.g., KCl, Na_2CO_3, $CuSO_4$, sugar, and alcohol are all soluble in water.
2. Non-polar substances are not soluble in polar solvent. Non polar covalent compounds are not soluble in water such as ether, benzene, and petrol are insoluble in water.
3. Non-polar covalent substances are soluble in non-polar solvents (mostly organic solvents). Grease, paints, naphthalene are soluble in ether or carbon tetrachloride etc.

(b) Solubility and solute- solvent interaction:

The solute-solvent interaction can be explained in terms of creation of attractive forces between the particles of solute and those of solvent. To dissolve one substance (solute) in another substance (solvent) following three events must occur.

(i) Solute particles must separate from each other
(ii) Solvent particles must separate to provide space for solute particles.
(iii) Solute and solvent particles must attract and mix up.

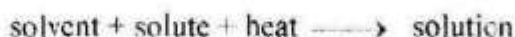
Solution formation depends upon the relative strength of attractive forces between solute- solute, solvent - solvent and solute - solvent. Generally solutes are solids. Ionic solids are arranged in such a regular pattern that the inter-ionic forces are at a maximum. If the new forces between solute and solvent particles overcome the solute-solute attractive forces, then solute dissolves and makes a solution. If forces between solute particles are strong enough than solute - solvent forces, solute remains insoluble and solution is not formed. Figure given below shows dissolution process by the interaction

of solvent molecules with the solute ions. The solvent molecules first pull apart the solute ions and then surround them. In this way solute dissolves and solution forms. For example, when NaCl is added in water it dissolves readily because the attractive interaction between the ions of NaCl and polar molecules of water are strong enough to overcome the attractive forces between Na^+ and Cl^- ions in solid NaCl crystal. In this process the positive end of the water dipole is oriented towards the Cl^- ions and the negative end of water dipole is oriented towards the Na^+ ions. These ion-dipole attractions between Na^+ ions and water molecules, Cl^- ions and water molecules are so strong that they pull these ions from their positions in the crystal and thus NaCl dissolves. It is shown in the given figure.

Interaction of solute and solvent to form solution

Q.7. Discuss the effect of temperature on solubility.

Ans. Effect of temperature on solubility: Temperature has major effect on the solubility of most of the substances. Generally it seems that solubility increases with the increase of temperature but it is not always true. When a solution is formed by adding a salt in solvent there are three possibilities with reference to effect of temperature on solubility. It is shown in the figure.

(i) Heat is absorbed: When salts like KNO_3, $NaNO_3$ and KCl are added in water, the test tube becomes cold. It means during dissolution of these salts heat is absorbed. Such process is called "endothermic".

$$\text{solvent} + \text{solute} + \text{heat} \longrightarrow \text{solution}$$

Solubility usually increases with the increase in temperature of such solutes. It means that heat is required to break the attractive forces between the ions of solute. This requirement is fulfilled by the surrounding molecules. As a result, their temperature falls down and test tube becomes cold.

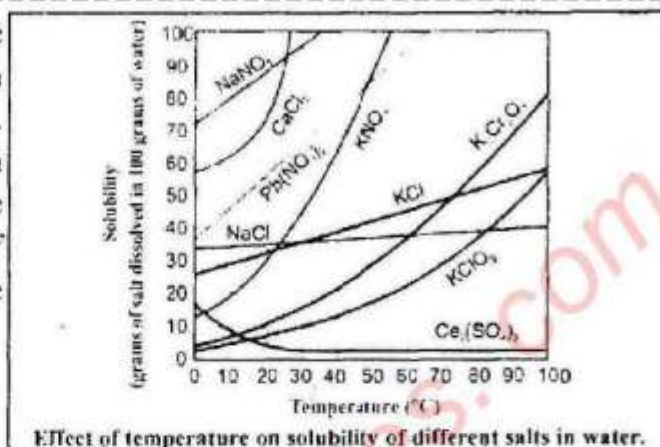
(ii) Heat is given out: On the other hand when salts like Li_2SO_4 and $Ce_2(SO_4)_3$ are dissolved in water, the test tube becomes warm i.e. heat is released during this dissolution.

$$\text{solvent} + \text{solute} \longrightarrow \text{solution} + \text{heat}$$

In such cases, the solubility of salt decreases with the increase of temperature. In such cases attractive forces among the solute particles are weaker and solute-solvent interactions are stronger. As a result, there is release of energy.

(iii) No change in heat: In some cases during a dissolution process neither the heat is

absorbed nor released. When salt like NaCl is added in water, the solution temperature remains almost the same. In such case temperature has a minimum effect on solubility figure given above shows the trend of solubilities of different salts with the increase in temperature.

Effect of temperature on solubility of different salts in water.

Test yourself 6.3

i. What will happen if the solute-solute forces are stronger than those of solute-solvent forces?

Ans. The solution will not formed. i.e. solute does not dissolve in solvent.

ii. When solute-solute forces are weaker than those of solute-solvent forces? Will solution form?

Ans. Yes, the solution will formed i.e. solute will dissolve in solvent.

iii. Why is iodine soluble in CCl_4 and not in water?

Ans. Iodine is a non-polar molecule which is only soluble in non-polar solvent (CCl_4) according to the principle like dissolve like.

iv. Why test tube becomes cold when KNO_3 is dissolved in water.

Ans. When KNO_3 is dissolved in water an endothermic process takes place hence tube becomes cold.

6.6 COMPARISON OF SOLUTION, SUSPENSION AND COLLOID

Q.8 Discuss the characteristics of solution, colloid and suspension. OR Compare the properties of solution colloid and suspension comparison of the characteristics of solution, colloidal and suspension

Solution	Colloid	Suspension
The particles exist in their simplest form i.e. as molecules or ions. Their diameter is 10^{-8} cm.	The particles are large consisting of many atoms; ions or molecules.	The particles are of largest size. They are larger than 10^{-5}cm in diameter.
Particles dissolve uniformly throughout and form a homogenous mixture.	A colloid appears to be a homogeneous but actually it is a heterogeneous mixture. Hence, they are not true solution. Particles do not settle down for a long time, therefore, colloids are quite stable.	Particles remain undissolved and form a heterogeneous mixture. Particles settle down after sometime.

Particles are so small that they can't be seen with naked eye.	Particles are large but can't be seen with naked eye.	Particles are big enough to be seen with naked eye.
Solute particles can pass easily through the filter paper.	Although particles are big but they can pass through a filter paper.	Solute particles cannot pass through filter paper.
Particles are so small that they cannot scatter the rays of light, thus do not show tyndall effect	Particles scatter the path of light rays thus emitting the beam of light i.e. exhibit the tyndall effect	Particles are so big that light is blocked and difficult to pass.

Test yourself 6.4

i. **What is difference between colloid and suspension?**

Ans.

Colloid	Suspension
(i) Particles can not be seen with naked eye. (ii) Particles can pass through the filter paper.	(i) Particles are big enough to be seen with naked eye. (ii) Particles can not pass through the filter paper.

ii. **Can colloids be separated by filtration, if not why?**

Ans. Colloid particles can not be separated by filtration because they are so small that they can pass through the filter paper.

iii. **Why are the colloids quite stable?**

Ans. Particles of colloids do not settle down for a long time hence colloids are quite stable.

iv. **Why does the colloid show tyndall effect?**

Ans. Colloids show tyndall effect because the particles of colloids scatter the path of light rays.

v. **What is tyndall effect and on what factors it depends?**

Ans. The scattering of the light by colloid particles is called tyndall effect. It depends upon the size of particles.

vi. **Identify as colloids or suspensions from the following: Paints, milk, milk of magnesia, soap solution.**

Ans.

Colloids	Suspension
Milk, soap solution	Milk of magnesia, Paints

vii. **How can you justify that milk is a colloid.**

Ans. Milk particles are so small that they do not settle down and passes through filter paper which is the character of colloid hence milk is colloid.

RELATIONSHIP OF SOLUTIONS TO DIFFERENT PRODUCTS IN THE COMMUNITY: Our body is made up of tissues, which are all composed of water based chemicals. The water becomes the best solvent in our body. We need an adequate supply of chemicals in the form of food, vitamins, hormones, and enzymes. For taking care of our health we need medicines. We find that chemicals and chemistry penetrate into every aspect of our life. Paper, sugar, starch, vegetable oils, ghee, essential oils, tannery, soap, cosmetics, rubber, dyes, plastics, petroleum, infact, there is almost nothing that we use in our daily life that is not a chemical. Some are usable as solid or gas but majority of them are used as solutions or suspensions.

Key Points

- Solution is a homogeneous mixture of two or more substances.
- Aqueous solution is formed by dissolving substances in water.
- The component which is lesser in quantity is called solute and the component in greater quantity is called solvent.
- A solution containing less amount of solute than that is required to saturate it at a given temperature is called unsaturated solution.
- A solution that is more concentrated than that of a saturated solution is called as supersaturated solution at that particular temperature.
- Solution may be dilute or concentrated depending upon the quantity of dissolved solute in solution.
- Concentration of solutions are expressed as % w/w, % w/v, % v/w and % v/v.
- The practical unit of concentration is molarity. It is the number of moles of solute dissolved in one dm^3 of solution.
- Solubility is defined as the number of grams of the solute dissolved in 100 g of solvent to prepare a saturated solution at a given temperature. It depends upon solute-solvent interactions and temperature.
- Colloidal solutions are false solutions and in these solutions particles are bigger than in the true solutions.

Exercise (Solved)

☆ **Multiple Choice Questions**

Put a (✓) on the correct answer.

1. **Mist is an example of solution:**
   (a) liquid in gas (b) gas in liquid (c) solid in gas (d) gas in solid
2. **Which one of the following is a liquid in solid solution?**
   (a) sugar in water (b) butter (c) opal (d) fog
3. **Concentration is ratio of:**
   (a) solvent to solute (b) solute to solution (c) solvent to solution (d) both a and b
4. **Which one of the following solutions contains more water?**
   (a) 2M (b) 1M (c) 0.5M (d) 0.25M
5. **A 5 percent (m/m) sugar solution means that:**
   (a) 5 g of sugar is dissolved in 90 g of water
   (b) 5 g os sugar is dissolved in 100 g of water
   (c) 5 g of sugar is dissolved in 105 g of water
   (d) 5 g of sugar is dissolved in 95 g of water
6. **If the solute-solute forces are strong enough than those of solute-solvent forces.**

The solute:
(a) dissolves readily (b) does not dissolve
(c) dissolves slowly (d) dissolves and precipitates

7. **Which one of the following will show negligible effect of temperature on its solubility?**
(a) KCl (b) KNO_3 (c) $NaNO_3$ (d) NaCl

8. **Which one of the following is heterogeneous mixture?**
(a) milk (b) ink (c) milk of magnesia (d) sugar solution

9. **Tyndall effect is shown by:**
(a) sugar solution (b) paints (c) jelly (d) chalk solution

10. **Tyndall effect is due to:**
(a) blockage of beam of light (b) non-scattering of beam of light
(c) scattering of beam of light (d) passing through beam of light

11. **If $10cm^3$ of alcohol is dissolved in 100 g of water, it is called:**
(a) % w/w (b) % w/v (c) % v/w (d) % v/v

12. **When a saturated solution is diluted it turns into:**
(a) supersaturated solution (b) saturated solution
(c) a concentrated solution (d) unsaturated solution

13. **Molarity is the number of moles of solute dissolved in:**
(a) 1 kg of solution (b) 100 g of slovent (c) $1dm^3$ of solvent (d) 1 dm^3 of solution

Answers:

1. liquid in gas **2.** butter **3.** solute to solution **4.** 0.25M
5. 5g of sugar is dissolved in 95g of water **6.** does not dissolve
7. NaCl **8.** milk of magnesia **9.** jelly
10. scattering of beam of light **11.** %v/w **12.** unsaturaged solution
13. 1 dm^3 of solution

☆Short Answer Questions.

1. **Why suspensions and solutions do not show tyndall effect, while colloids do?**

Ans. Solutions do not show tyndall effect because their particles are so small that they can not scatter the rays of light.

Suspensions do not show tyndall effect because their particles are so big that light is blocked and difficult to pass.

Colloid show tyndall effect because their particles are big enough to scatter the beam of light.

2. **What is the reason for the difference between solutions, colloids and suspensions?**

Ans. The particles size is the basic reason for the difference between solutions, colloids and suspension.

3. Why the suspension does not form a homogeneous mixture?

Ans. Suspension do not form a homogeneous mixture because its particles remain undissolved and settle down after sometimes.

4. How will you test whether given solution is a colloidal solution or not?

Ans. Colloidal solution can be checked by tyndall effect.
If the solution scatter the beam of light, then it is a colloidal solution otherwise not.

5. Classify the following into true solution and colloidal solution:
Blood, starch solution, glucose solution, toothpaste, copper sulphate solution, silver nitrate solution.

Ans.

True solution	Colloidal solution
Glucose solution, copper sulphate solution silver nitrate solution	Toothpaste, starch solution, blood

6. Why do we stir paints thoroughly before using?

Ans. Paint is a heterogeneous mixture if we not stir the point before use, the particles will settle down.

7. Which of the following will scatter light and why?
sugar solution, soap solution and milk of magnesia.

Ans. Soap solution will scatter light because it shows tyndall effect.

8. What do you mean, like dissolves like? Explain with examples.

Ans. For answer see Q. No. 6 (a)

9. How does nature of attractive forces of solute-solute and solvent-solvent affect the solubility?

Ans. For answer see Q. No. 6 (b)

10. How can you explain the solute-solvent interaction to prepare a NaCl solution?

Ans. For answer see Q. No. 6 (b)

11. Justify with an example that solubility of a salt increases with the increase in temperature.

Ans. For answer see Q. No. 7

12. What do you mean by volume/volume %?

Ans. For answer see Q. No. 3(b)

Long Answer Questions

Q.1 What is saturated solution and how it is prepared?

Ans. For answer see Q. No. 1

Q.2 Differentiate between dilute and concentrated solutions with a common example.

Ans. For answer see Q. No. 1

Q.3 Explain, how are dilute solutions prepared from concentrated solutions?

Ans. For answer see Q. No. 5

Q.4 What is molarity and give its formula to prepare molar solution?

Ans. For answer see Q. No. 4(a)

Q.5 Explain the solute-solvent interaction for the preparation of solution.

Ans. For answer see Q. No. 6(a)

Q.6 What is general principle of solubility?

Ans. For answer see Q. No. 6(a)

Q.7 Discuss the effect of temperature on solubility.

Ans. For answer see Q. No. 7

Q.8 Give five characteristics of colloids.

Ans. For answer see Q. No. 8

Q.9 Give at least five characteristics of suspension.

Ans. For answer see Q. No. 8

Numericals

Q.1 A solution contains 50 g of sugar dissolved in 450 g of water. What is concentration of this solution?

Solution: Mass of solute (sugar) = 50g

Mass of water (solvent) = 450g

Mass of solution = 50 + 450

= 500g

Concentration of solution (m/m) = ?

As

Concentration of solution (m/m) $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

$\therefore$ Concentration of solution $= \frac{50}{500} \times 100 = 10\%$ (m/m)

Result: Concentration of solution = 10%

Q.2 If 60cm³ of alcohol is dissolved in 940cm³ of water, what is concentration of this solution?

Solution: Volume of solute (alcohol) = $60cm^3$

Volume of solvent (Water) = $940cm^3$

Volume of solution [solute + solvent] $= 940 + 60$

$= 1000\ cm^3$

As

Concentration of solution (v/v) $= \dfrac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$

∴ Concentration of solution $= \dfrac{60}{1000} \times 100$

$= 6\% = \mathbf{6\ \%\ v/v}$

Result: Concentration of solution = 6% v/v

Q.3 How much salt will be required to prepare following solutions (atomic mass: K = 39; Na = 23; S = 32; O = 16 and H = 1)

a. 250cm³ of KOH solution of 0.5M

Solution:

Amount of solute (KOH) in gram= ?

Volume of solution $= 250cm^3 = \dfrac{250}{1000}$

$= 0.25dm^3$

Molarity of solution $= M = 0.5M$

Formula mass of solute, KOH $= 39 + 16 + 1$

$= 56$

As

$$\text{Molarity} = \frac{\text{Mass of solute in grams}}{\text{Formula mass of solute} \times \text{Volume of solution in } dm^3}$$

By putting the values;

$$0.5 = \frac{\text{Mass of solute in grams}}{56 \times 0.25}$$

∴ Mass of solute in grams $= 0.5 \times 56 \times 0.25 = 7$

Result: Amount of solute (KOH) = 7g

b. 600cm³ of $NaNO_3$ solution of 0.25M

Solution: Amount of solute ($NaNO_3$) = ?

Volume of solution in dm^3 $= \dfrac{600}{1000} = 0.6\ dm^3$

Molarity of solution $= M = 0.25M$

Molar mass of $NaNO_3$ $= (23\times1) + (14\times1) + (16\times3)$

$= 23 + 14 + 48 = 85$ g/mole

As

$$\text{Molarity} = \frac{\text{Mass of solute in grams}}{\text{Molar mass} \times \text{Volume of solution in } dm^3}$$

∴ By putting the values;

$$0.25 = \frac{\text{Mass of solute } (NaNO_3) \text{ in gram}}{85 \times 0.6}$$

$0.25 \times 85 \times 0.6$ = Mass of solute

∴ Mass of solute $= 12.75g$

Result: Amount of $NaNO_3 = 12.75g$

c. ***800cm³ of Na_2SO_4 solution of 1.0M***

Solution: Amount of Na_2SO_4 in grams = ?

Volume of solution in dm³ $= \frac{800}{1000} = 0.8\ dm^3$

Molarity of solution $= M = 1.0\ M$

Molar mass of Na_2SO_4 $= (23\times2) + (32\times1) + (16\times4)$

As $= 46 + 32 + 64$

$= 142$

$$\text{Molarity} = \frac{\text{Amount of solute in grams}}{\text{Molar mass of solute} \times \text{Volume of solution in dm}^3}$$

∴ By putting the values;

$$1 = \frac{\text{Amount of solute}}{142 \times 0.8}$$

$1 \times 142 \times 0.8$ = Amount of solute

∴ Amount of solute $= 113.6$

Result: Amount of solute $(Na_2SO_4) = 113.6g$

Q.4 When we dissolve 20 g of NaCl in 400 cm³ of solution, what will be its molarity?

Solution: Amount of solute (NaCl) in grams $= 20g$

Volume of solution in dm³ $= \frac{400}{1000}$

$= 0.4dm^3$

Molarity of solution $= M = ?$

Molar mass of solute, NaCl $= 23+35.5 = 58.5g$

As

$$\text{Molarity} = \frac{\text{Amount of solute in grams}}{\text{Molar mass of solute} \times \text{Volume of solution in dm}^3}$$

By putting the values;

$$\text{Molarity} = \frac{20}{58.5 \times 0.4} = \frac{20}{23.4} = 0.85$$

Result: Molarity of solution $= M = 0.85M$

Q.5 We desire to prepare $100cm^3$ 0.4M solution of $MgCl_2$, how much $MgCl_2$ is needed?

Solution: Amount of solute ($MgCl_2$) = ?

Volume of solution in dm^3 $= \frac{100}{1000} = 0.1dm^3$

Molarity of solution = M = 0.4M

Molar mass of solute, $MgCl_2$ = (24×1) + (35.5×2)

As = 95

$$\text{Molarity} = \frac{\text{Amount of solute } (MgCl_2) \text{ in grams}}{\text{Molar mass of solute} \times \text{Volume of solution in } dm^3}$$

= By putting the values:

$$0.4 = \frac{\text{Amount of solute}}{95 \times 0.1}$$

Amount of $MgCl_2$ = 0.4×95×0.1

= 3.8g

Result: Amount of $MgCl_2$ = 3.8g

Q.6 12M H_2SO_4 solution is available in the laboratory. We need only $500cm^3$ of 0.1 M solution, how it will be prepared?

Solution: Given H_2SO_4 = Required H_2SO_4

$$M_1V_1 = M_2V_2$$

$$12 \times V_1 = 0.1 \times 500$$

$$\therefore V_1 = \frac{0.1 \times 500}{12}$$

$$= 4.166 \text{ cm}^3$$

Result: Take $4.166cm^3$ of 0.1M H_2SO_4 in a measuring flask of $500cm^3$. and mark the volume upto mark by adding distilled water.

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

6.1	Solution
6.2	Saturated Solution
6.3	Types of Solution

☆ Tick the correct answer.

1. **The component of solution which is present in smaller quantity is called:** (FBD. GI)
   (A) Solvent (B) Saturated solution
   (C) Solute (D) Unsaturated solution
2. **___________ is heterogeneous mixture.** (MLN. GI, BWP. GI, GRW. GI)
   (A) Milk (B) Ink (C) Milk of Magnesia (D) Sugar Solution
3. **Air is an example of solution:** (SGD. GII, BWP. GII)
   (A) gas in gas (B) gas in solid (C) solid in gas (D) gas in liquid
4. **A universal solvent on earth is:** (SWL. GI)
   (A) water (B) alcohol (C) ammonia (D) ether
5. **Metal alloys is an example of:** (LHR. GI, MLN. GII)
   (A) Liquid in gas (B) Gas in liquid (C) Solid in gas (D) Solid in solid
6. **Example of liquid in liquid is:** (GRW. GII, SGD. GII)
   (A) alcohol in water (B) butter in water (C) fog (D) mist
7. **Brass is a solid solution of:** (FBD. GII)
   (A) Cu + Zn (B) Cu + Ni (C) Cu + Fe (D) Cu + Na
8. **Mist is an example of solution:** (SWL. GII, DGK. GII, BWP. GII, FBD. GI)
   (A) liquid in gas (B) gas in liquid (C) solid in gas (D) liquid in solid
9. **Fog is an example of a solution of:** (RWP. GII)
   (A) Liquid in gas (B) Gas in liquid (C) Solid in gas (D) Gas in solid
10. **Which one the following is a liquid in solid solution:** (LHR. GI)
    (A) Sugar in water (B) Butter (C) Alcohol in water (D) Fog
11. **Metals alloy are:** (LHR. GII)
    (A) Solution of solid in gass (B) Solution of solid in liquid
    (C) Solution of solid in solid (D) Solution of gas in soild

12. An example of true solution is: (FBD. GII)
(A) Starch solution (B) Tooth paste (C) Soap solution (D) ink in water

13. Opal is an example of solution: (RWP. GI)
(A) liquid in gass (B) solid in gas (C) solid in solid (D) gas in solid

14. The types of solution are: (RWP. GII)
(A) 8 (B) 7 (C) 9 (D) 10

Answers

1.	Solute	2.	Milk of Magnesia	3.	gas in gas	4.	water
5.	Solid in solid	6.	alcohol in water	7.	Cu + Zn	8.	liquid in gas
9.	Liquid in gas	10.	Butter	11.	Solution of solid in solid		
12.	ink in water	13.	solid in solid	14.	9		

☆ **Give short answer to the following questions.**

1. Define solvent and solute. (LHR. GI, BWP. GII, MLN. GI, SGD. GI, RWP. GI, FBD. GI)

Ans. **Solute:** The component of solution present in relatively lesser amount is called solute.

Solvent: The component of solution present in relatively large amount is called solvent. Solvent always dissolves solutes.

2. Define aqueous solutions. (LHR. GII, MLN. GI, FBD. GI, SWL. GI, MLN. GII, BWP. GII)

Ans. The solution which is formed by dissolving substance in water is called aqueous solution. In aqueous solution, water is in large amount.

3. Differentiate between solution and aqueous solution. (FBD. GII, RWP. GII)

Ans.

Solution	Aqueous Solution
The homogeneous mixture of two or more things e.g. solution fo sugar.	A type of solution formed by dissolving something in water. e.g. solution of sugar and salt in water, is aqueous solution.

4. Why water is called a universal solvent? (SWL. GI)

Ans. Because mostly compounds in the universe dissolve in water. That is why water is known as universal solvent.

5. What is the difference between solution and mixture? (BWP. GII)

Ans. The homogeneous mixture of different substances is known as solution. The component in solution are in same phase, while the component of mixture are not uniform. They are in different phases.

6. How can you distinguish between solution and pure solvent? (LHR. GI)

Ans. Through evaporation, difference between solution and pure liquid is found.

7. What is homogeneous mixture of two or more substances? (GRW. GII)

Ans. The homogeneous mixture of two or more thing is called solution.

8. What do you mean by unsaturated solution? (LHR. GI, DGK. GI, SGD. GII)

Ans. A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution.

9. Define saturated solution. (GRW. GII)

Ans. A solution contain maximum amount of solute at a given temperature is called saturated solution. At particle level, the saturated solution is that in which undissolved solute is in equilibrium with dissolved solute.

10. Differentiate between Saturated and Unsaturated Solutions. (MLN. GII)

Ans. **Saturated solution:** A solution containing maximum amount of solute at a given temperature called saturated solution.

Unsaturated solution: A solution containing lesser amount of solute than that which is required to saturate it at a given temperature.

11. Define super saturated solution. (SGD. GII, DGK. GI)

Ans. A solution which required more amount of solute than that required for preparing it saturated solution at particular temperature is called super saturated solution. These are more concentrated than saturated solution.

12. What is alloy? Give an example. (GRW. GI, SWL. GI, MLN. GII)

Ans. Mixture of two metals is called alloy. For example: Brass is a mixture of copper and zinc.

13. Write down example of a solution in which solute is liquid and solvent is gas. (LHR. GII)

Ans. Fog, Mist

14. What is solid-liquid solution? Give an example. (SGD. GII)

Ans. When solid (solute) is dissolved in liquid (solvent) it is known as solid liquid solution. e.g. sugar in water.

15. Write down two examples of liquid in liquid solution. (DGK. GI)

Ans.
1. The solution of water in alcohol.
2. The solution of benzene in toluene.

6.4	Concentration Units
6.5	Solubility
6.6	Comparison of Solution, Suspension and Colloid

☆ Tick the correct answer.

1. Which one of the following solutions has less water? (GRW. GI, DGK. GI)

(A) 0.25M (B) 0.50M (C) 0.60M (D) 2.0M

2. **If 10cm³ of alcohol is dissolved in 100 g of water, it is called:** (SWL. GI, DGK. GI)
(A) $\%\frac{w}{w}$ (B) $\%\frac{w}{v}$ (C) $\%\frac{v}{w}$ (D) $\%\frac{v}{v}$

3. **Which solution contains more water?** (RWP. GII, SGD. GI)
(A) 2M (B) 1M (C) 0.5M (D) 0.25M

4. **The concentrated solution of common salt in water is called:** (MLN. GI)
(A) Brine (B) Benzene (C) Alcohol (D) Toluene

5. **Which one of the following will show negligible effect of temperature on its solubility:** (LHR. GII, RWP. GI, DGK. GII)
(A) KCl (B) KNO_3 (C) $NaCl$ (D) $NaNO_3$

6. **Paints and ether are mixable because:** (GRW. GII)
(A) both are polar (B) both are non-polar
(C) paints are polar but ether is non-polar
(D) both have different chemical nature

7. **Number of moles of solute in one dm^3 of the solution is called:** (SGD. GII)
(A) solubility (B) molarity (C) colloid (D) suspension

8. **Tyndall effect is due to __________ beams of light.** (RWP. GI)
(A) blockage of (B) non-scattering (C) passing through (D) scattering of

9. **Which one of the following exhibits the tyndall effect:** (SWL. GII)
(A) solution (B) colloid (C) suspension (D) solvent

Answers

1. 2.0M	**2.** $\%\frac{v}{w}$	**3.** 0.25M	**4.** Brine
5. $NaCl$	**6.** both are non-polar	**7.** molarity	**8.** scattering of
9. colloid			

☆ Give short answer to the following questions.

1. **What is difference between dilute solution and concentrated solution?**
(LHR. GI, SWL. GII, MLN. GI, GRW. GI, SGD. GI, BWP. GII, MLN. GII)

Ans. **Dilute Solution:** A type of solution in which more solvent is present than solute.
Concentrated Solution: A solution in which amount of solute is more than solvent.

2. **What is % v/m?** (LHR. GI, MLN. GII, SWL. GI, DGK. GI, BWP. GI)

Ans. **Percentage v/m:** It is the volume in cm^3 of a solute dissolved in 100g of solution. For example 10% v/m alcohol solution in water means $10cm^3$ of alcohol is dissolved in water to make 100g solution. In such solution the mass of solution is under consideration, volume is not considered.

$$\%v/m = \frac{\text{Volume of solute } (cm^3)}{\text{Mass of solution}} \times 100$$

3. Define molarity. (LHR. GI, DGK. GI, FBD. GI, LHR. GII, SGD. GII)

Ans. It is defined as number of moles of solute present in one dm^3 of the solution. It is represented by M. Molarity is a concentration unit.

$$\text{Number of moles} = \frac{\text{mass of solute in grams}}{\text{molar mass of solute}}$$

$$M = \frac{\text{mass of solute in gram}}{(\text{molar mass of solute } gmol^{-1}) \times (\text{volume of solution in } dm^3)}$$

4. What is percentage $\left(\%\frac{m}{m}\right)$? (GRW. GI, SGD. GI, RWP. GII)

Ans. It is number of grams of solute dissolved in 100 grams of solution. For example 10% m/m sugar solution means that 10g of sugar is dissolved in 90gm of water to make the solution 100g.

5. What is meant by $\left(\%\frac{m}{v}\right)$? (GRW. GII, SGD. GII, BWP. GII)

Ans. It is number of gram of solute dissolved in $100cm^3$ of solution. For example 10% (m/v) sugar solution contains 10g of sugar is dissolved in $100cm^3$ of solution. The exact volume of solvent is not mentioned.

$$\%\frac{\text{mass}}{\text{volume}} = \frac{\text{mass of solute}}{\text{volume of solution}} \times 100$$

6. What do you mean by $^{Volume}/_{Volume}$ %? (DGK. GI, FBD. GII)

Ans. The amount of volume of solute in cm^3, which is dissolved in $100cm^3$ of solution, called percentage v/v.

$$\text{percentage } {}^{v}/_{v} = \frac{\text{volume of solute } (cm^3)}{\text{volume of solution } (cm^3)}$$

7. How one molar solution is prepared? (LHR. GII)

Ans. To prepare 1 molar solution, one mole of solute is dissolved in water to make volume of solution $1dm^3$. This solution is make in measuring flask.

8. How one molar solution of NaOH is prepared? (GRW. GI, MLN. GI)

Ans. To prepare the 1 molar of sodium hydroxide, take 1 mole (40g) of salt and dissolve in water to make the volume of $1dm^3$.

9. Write the names of ways in which concentration units are expressed. (GRW. GI)

Ans. Percentage mass / mass (% m/m)

Percentage mass / volume (% m/v)

Percentage volume / mass (% v/m)

Percentage volume / volume (% v/v)

10. How much NaOH is required to prepare its 500cm^3 of 0.4 M solution? (GRW. GII)

Ans. Molar mass of NaOH = 40gmol^{-1}

Volum in dm^3 = $\frac{500cm^3}{1000cm^3} \times 1dm^3 = 0.5dm^3$

Molarity = $\frac{\text{Mass of solute in gram}}{\text{molar mass(gmol}^{-1}\text{)} \times \text{volume of solution(dm}^3\text{)}}$

mass of solute = Molarity×molar mass of solute × volume of solution

= 0.4 × 40 × 0.5

= 8g

11. Which one of the following solution is more concentrated one molar solution or 3 molar? Why? (SWL. GI)

Ans. 3 molar solution is more concentrated, because it has dissolved large amount of solute.

12. How much volume of 0.1M solution is required if you are asked to prepare a solution of 0.1 molar having volume 100cm^3? (SWL. GII)

Ans.

$M_1 = 0.1$

$V_1 = ?$

$M_2 = 0.01$

$V_2 = 100cm^3$

$M_1V_1 = M_2V_2$

$0.1V_1 = 0.01 \times 100$

$V_1 = \frac{0.01 \times 100}{0.1} = 10cm^3$

13. Define the term dilute solution. (RWP. GI)

Ans. A type of solution, in which the dissolved amount of solute is low, is known as dilute solution.

14. How can you prepare 1dm^3 solution of NaOH having 0.5M molarity? (RWP. GI)

Ans. Molar mass of NaOH = 40

Mass of solute = Molarity×molar mass of solute×volume of solution

= 0.5 × 40 × 1

= 20g

20 gram of NaOH is dissolved in sufficient solution, that its volume become 1dm^3.

15. 5cm³ acetone is dissolved to prepare 90cm³ aqueous solution. Calculate the percentage v/v of the solution. (DGK. GII, BWP. GI)

Ans. Concentration of solution v/v $= \dfrac{\text{volume of solute}}{\text{volume of solution}} \times 100$

$$= \frac{5}{90} \times 100 = 5.5$$

16. Define solubility. (GRW. G II, MLN. GI, BWP. GII, RWP. GI, GRW. GII)

Ans. Solubility is defined as number of grams of solute dissolved in 100 gram of solvent to prepare saturated solution at particular temperature.

17. Why test tube becomes cold when KNO_3 is dissolved in water? (SWL. GII, RWP. GII, FBD. GI)

Ans. When KNO_3 is dissolved in water, an endothermic process has been taking place, that's why test tube becomes cold.

18. What is general principle of solubility? (SGD. GI)

Ans. The common rule of solubility is "like dissolve like" polar compounds are dissolved in polar solvents. Non-polar compounds do not dissolved in polar solvents. Non polar covalent compound dissolved in non polar solvents.

19. What is the effect of temperature on solubility? (RWP. GII, FBD. GI)

Ans. Temperature has greater effect on solubility usually, by increasing temperature, solubility is also increased.

When a solution is made by adding some salt in a solvent, there will be three conditions.

i. Heat is absorbed
ii. Heat is released
iii. There is no change in heat.

20. Justify with example that solubility of a salt increases with increase in temperature. (FBD. GII)

Ans. If during salt dissolving process, heat is absorbed, then the solubility of such salts increases with increase of temperature. e.g when KNO_3 is dissolved in water an endothermic process takes place hence tube become cold.

21. What is difference between true solution and colloidal solution? (LHR. GII)

Ans.

True Solution	Colloidal Solution
The particles of true solution are very small, that they can't scatter the rays of light.	The particles of colloidal are big enough, that scatter the rays of light, when light passes through them.

22. What is tyndall effect? (GRW. GI, FBD. GII, RWP. GII, DGK. GII)

Ans. When light is passed through colloid, the particles of colloid scatter the light rays.

This phenomenon is called tyndall effect. It depends upon size of particles (solute).

23. How can you justify that milk is a colloid? (FBD. GII)

Ans. Milk is colloid because it's particles pass through filter paper. It's particles scatter the light rays and show tyndall effect.

24. Define suspension with an example. (FBD. GII, SWL. GII)

Ans. "Heterogeneous mixture of undissolved particles of solute in a medium called suspension". The particles of suspension are big enough to be seen with naked eye. e.g. chalk solution and paints.

25. Write two properties of suspensions. (SGD. GI)

Ans. The particles of solute in a suspension are big enough to be seen with naked eye. In suspension the solute particles cannot pass through filter paper.

26. On what factors tyndall effect depends? (SGD. GI)

Ans. It depend upon the wavelength of light and size of solute particles.

27. Write down any four examples of colloidal solution. (RWP. GI)

Ans. Milk, Ink, Jelly, toothpaste are the examples of colloidal solution.

28. Why do true solutions not show Tyndall effect? (RWP. GI)

Ans. True solutions do not show the phenomenon of tyndal effect, because the particles of true solution can not scatter the light rays, when light passing through them.

29. Why does not the Suspension form Homogenous Mixture? (BWP. GI, SGD. GII)

Ans. The particles of suspension are big enough to be seen with naked eye. They do not dissolved in solvent and easily settled down, that is why suspension, do not form homogenous mixture.

30. Why we stir paints thoroughly before using? (BWP. GI. DGK. GII)

Ans. Paints are suspension, their heavy particles are settled down, that is why paints are stirrer before use.

31. How will you test whether given solution is a Colloidal Solution or not? (BWP. G I)

Ans. Through the phenomenon of tyndal effect, it has to be decided that either given solution is colloid or not.

32. Why does the colloids show tyndal effects? (BWP. GII, GRW. GII)

Ans. The particles of colloids are big, but not enough to be seen with naked eye. When light passes through them. they scatter the light rays. This phenomenon of scattering is known as tyndall effects.

33. Give two examples of colloids. (LHR. GI)

Ans. Blood, Jelly and tooth paste are examples of colloids.

34. Define colloid. (LHR. GII, FBD. GI)

Ans. Colloids are the solution, in which particles are larger than the particles of solute of

real solution, but not big enough to be seen with naked eye.

35. How colloids differ from solutions? (GRW. GII)

Ans. The component of solution are very small, that they do not scatter light, when light passes through them. While the particles of colloids are big enough to scatter the light, when light is passing through them.

36. Why are the colloids quite stable? (FBD. GII)

Ans. In colloids, particles are dissolved, but do not settled down for a long time. That is why colloids are quite stable.

37. Write any two differences between solution and suspension. (MLN. GI, BWP. GI)

Ans. Solution:

1. A homogeneous mixture of two or more component.
2. The particles of solution can not be seen with naked eye.

Suspension:

1. The heterogeneous mixture of undissolved particles.
2. The particles are big enough to be seen with naked eye.

38. Identify as colloids or suspensions from the following. Chalk, Soap Solution, Milk, Paints (SWL. GI)

Ans. Suspension = Paints and chalk

Colloids = Milk and Soap solution

39. Differentiate between colloid and suspension. (SWL. GI, SGD. GII, SGD. GI)

Ans.

Colloids	Suspension
1- The particles of colloids show tyndal effect.	1- The particles of suspension do not show tyndal effect.
2- The particles of colloid can pass through filter paper.	2- Its particles do not pass through filter paper.
3- Its particles do not settled down for a long time.	3- It particles settled down after some time.

ELECTROCHEMISTRY

Major Concepts:

7.1 Oxidation and reduction
7.2 Oxidation states and rules for assigning oxidation states
7.3 Oxidizing and reducing agents
7.4 Oxidation - reduction reactions
7.5 Electrochemical cells
7.6 Electrochemical industries
7.7 Corrosion and its prevention

Time allocation

Teaching periods	18
Assessment periods	03
Weightage	18%

Students Learning Outcomes:

Students will be able to:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- Identify the oxidizing and reducing agents in a redox reaction.
- Define oxidation state.
- State the common rules used for assigning oxidation numbers to free elements, ions (simple and complex), molecules, atoms.
- Determine the oxidation number of an atom of any element in a compound.
- Describe the nature of electrochemical processes.
- Sketch an electrolytic cell, label the cathode and the anode.
- Identify the direction of movement of cations and anions towards respective electrodes.
- List the possible uses of an electrolytic cell.
- Sketch a Daniel cell, labelling the cathode, the anode, and the direction of flow of the electrons.
- Describe how a battery produces electrical energy.
- Identify the half-cell in which oxidation occurs and the half-cell in which reduction occurs given a voltaic cell.
- Distinguish between electrolytic and voltaic cells.
- Describe the methods of preparation of alkali metals.
- Describe the manufacture of sodium metal from fused NaCl.
- Identify the formation of by products in the manufacture of sodium metal from fused NaCl.

- Describe the method of recovering metal from its ore.
- Explain electrolytic refining of copper.
- Define corrosion.
- Describe rusting of iron as an example of corrosion.
- Summarize the methods used to prevent corrosion.
- Explain electroplating of metal on steel (using examples of zinc, tin and chromium plating).

7.1 OXIDATION AND REDUCTION REACTIONS

Q.1. Define the following.

(i) Electrochemistry ***(ii) Non-spontaneous reactions***
(iii) Spontaneous reaction ***(iv) Redox reaction***
(v) Oxidation ***(vi) Reduction***

Ans.(i) Electrochemistry: The branch of chemistry which deals with the conversion of chemical energy into electrical energy and electrical energy into chemical energy.

It deals with the relationship between electricity and chemical reactions.

(ii) Non- Spontaneous reactions: The reactions which can not take place by their own are called non-spontaneous reactions.

(iii) Spontaneous reactions: The reactions which can take place by their own without any external agent are called spontaneous reaction.

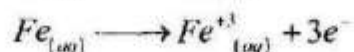
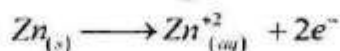
(iv) Redox reactions: The reactions in which oxidation as well as reduction take place are called redox reactions e.g.

$$CuO + H_2 \longrightarrow Cu + H_2O$$

(v) Oxidation: The oxidation can se defined in following ways.

1. Addition of oxygen is called oxidation. e.g. $C + O_2 \longrightarrow CO_2$
2. Removal of hydrogen is called oxidation. e.g. $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$
3. Loss of electrons is called oxidation. e.g.

$$Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$$

$$Fe_{(aq)} \longrightarrow Fe^{+3}_{(aq)} + 3e^-$$

(vi) Reduction: The reduction can be defined in following ways.

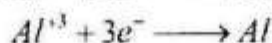
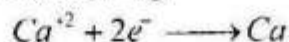
1. Addition of hydrogen is called reduction. e.g.

$$Cl_2 + H_2 \longrightarrow 2HCl$$

2. Removal of oxygen is called reduction. e.g.

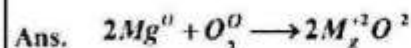
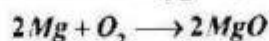
$$CuO + H_2 \longrightarrow Cu + H_2O$$

3. Gain of electrons is called reduction e.g.

$$Ca^{+2} + 2e^{-} \longrightarrow Ca$$
$$Al^{+3} + 3e^{-} \longrightarrow Al$$

Test yourself 7.1:

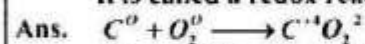
i. **How can you justify that a reaction between magnesium and oxygen is a redox reaction, while the reaction shows only addition of oxygen (oxidation)**

$$2Mg + O_2 \longrightarrow 2MgO$$

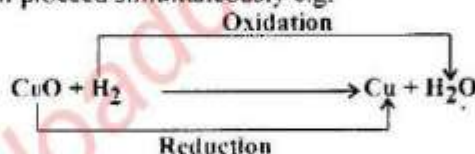
Ans. $2Mg^{0} + O_2^{0} \longrightarrow 2Mg^{+2}O^{-2}$

The oxidation state of "Mg" increases from zero to "+2" hence it is oxidation. While the oxidation state of oxygen decreases from zero to, "-2" hence it is reduction.

The above reaction is redox because both oxidation and reduction reactions are taking place.

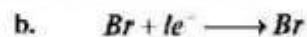
ii. **A reaction between carbon and oxygen involved only addition of oxygen (oxidation), but, it is called a redox reaction. comment on this.**

Ans. $C^{0} + O_2^{0} \longrightarrow C^{+4}O_2^{-2}$

The oxidation state of carbon increases from zero to "+4" hence it is oxidation while the oxidation state of oxygen decreases from zero to, "-2" hence it is reduction. The above reaction is redox reaction because both oxidations and reductions are taking place.

iii. **Oxidation and reduction proceed simultaneously. Explain, with an example.**

Ans. Oxidation and reduction proceed simultaneously e.g.

iv. **Identify which of the following is oxidation or reduction reaction**

a. $K \longrightarrow K^{+} + 1e^{-}$ b. $Br + 1e^{-} \longrightarrow Br^{-}$

c. $Cu \longrightarrow Cu^{+2} + 2e^{-}$ d. $I^{-} \longrightarrow I + 1e^{-}$

e. $Fe^{+2} \longrightarrow Fe^{+3} + 1e^{-}$

Ans. (a) $K \longrightarrow K^{+} + 1e^{-}$
Oxidation reaction

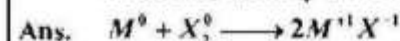
(b) $Br + 1e^{-} \longrightarrow Br^{-}$
Reduction reaction

(c) $Cu \longrightarrow Cu^{+2} + 2e^{-}$
Oxidation reaction

(d) $I^{-} \longrightarrow I + 1e^{-}$
Oxidation reaction

(e) $Fe^{+2} \longrightarrow Fe^{+3} + 1e$
Oxidation reaction

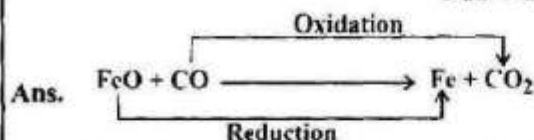
v. **An element M reacts with another element X to form MX_2. In terms of loss or gain of electrons, identify the element which is oxidized and which is reduced.**

Ans. $M^{0} + X_2^{0} \longrightarrow 2M^{+1}X^{-1}$

"M" is oxidised because its oxidation state increases from "0" to "+1".

X_2 reduced because its oxidation state decreases from '0' to '-1'.

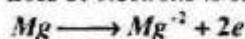
vi. How can you justify that the following reaction is not only an oxidation reaction but also a complete redox reaction.

$$FeO + CO \longrightarrow Fe + CO_2$$

Ans. $FeO + CO \longrightarrow Fe + CO_2$ (Oxidation: CO → CO_2; Reduction: FeO → Fe)

The above reaction is redox reaction reaction because FeO, reduced to "Fe" by removal of oxygen while "CO" is oxidized to CO_2 by giving of oxygen.

vii. Explain the term oxidation on the basis of electronic concept with an example.

Ans. Loss of electrons is called oxidation. e.g.

$$Mg \longrightarrow Mg^{+2} + 2e$$

7.2 OXIDATION STATE AND RULES FOR ASSIGNING OXIDATION STATE

Q.2. Define oxidation state. (Oxidation number). Discuss the rules of assigning oxidation number (oxidation state).

Ans. Oxidation state / oxidation number: The number which represents an apparent charge [positive, negative or zero] which an atom of an element would have in a molecule or ion is called oxidation number or oxidation state.

Rules for assigning oxidation numbers (O.N.):

1. The oxidation number of all elements in the free state is zero.
2. The oxidation number of an ion consisting of a single element is the same as the charge on the ion.
3. The oxidation number of different elements in the periodic table is: in Group 1 it is + 1, in Group- II it is +2 and in Group -III it is +3.
4. The oxidation number of hydrogen in all its compounds is +1. But in metal hydrides it is -1.
5. The oxidation number of oxygen in all its compounds is -2. But it is -1 in peroxides and +2 in OF_2.
6. In any substance the more electronegative atom has the negative oxidation number.
7. In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.
8. In ions, the algebraic sum of oxidation number equals the charge on the ion.

Remember!

It is important to note that while assigning oxidation numbers the sign precedes the number. It is written as +2. whereas, the apparent charge on an atom, ion or molecule which is called valency, is written as the sign followed by the number i.e. 2+.

Example 7.1:

Find oxidation number of nitrogen in HNO_3 when the oxidation numbers of H = +1 and O = -2

Solution: By applying formula in compound sum of all oxidation numbers is zero. In

case of this compound HNO_3 it becomes.

[O.N. of H] + [O.N. of N] + 3[O.N. Of O] = 0

Putting the values in above formula

[+1] + [O.N. of N] + 3[-2] = 0

+1 + O.N. of N + [-6] = 0

or O.N. of Nitrogen = 6–1

= + 5

Example: 7.2:

Calculate the oxidation number of sulphur in H_2SO_4. When O.N. of H = +1 and O.N. of O = -2

Solution:

Applying the formula of H_2SO_4.

2[O.N. of H] + [O.N. of S] + 4[O. N. Of O] = 0

Putting the values in above formula

2[+1] + [O.N. of 5] + 4 [-2] = 0

2 + [O.N. of S] - 8 = 0

O.N. of S **= 8 –2 = +6**

Example 7.3:

Find out the oxidation number of chlorine in $KClO_3$.

As O.N. of K = +1 and O.N. of O = -2

Solution:

Putting the values in formula, we get

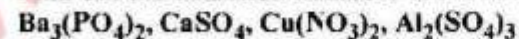
[+1] + [O.N. of Cl] + 3[-2] = 0

[+1] + [O.N. Cl] + [-6] = 0

O. N. of Cl = 6 – 1

= + 5

Test yourself 7.2:

i. **Find out the oxidation numbers of the following elements marked in bold in the formulae: $Ba_3(PO_4)_2$, $CaSO_4$, $Cu(NO_3)_2$, $Al_2(SO_4)_3$**

Ans. $Ba_3(PO_4)_2$

(+2)3 + (P)2 + (-2) 8 = 0

+6+2P-16 = 0

2P-10 = 0

2P = 10

$P = \frac{10}{2} = +5$

O.N of P = +5

$CaSO_4$

(+2)1+(S)1+(-2)4 = 0

$+2+S-8 = 0$

$S - 6 = 0$

O.N of S = +6

$Cu(NO_3)_2$

$(+2)1+(N)2+(-2)6 = 0$

$+2+2N-12 = 0$

$2N = 10$

O.N of N = $\frac{10}{2} = +5$

$Al_2(SO_4)_3$

$(+3)2+(S)3+(-2)12 = 0$

$+6+3S-24 = 0$

$3S - 18 = 0$

$S = 18/3 = +6$

O.N of S = +6

ii. In a compound MX_3, find out the oxidation number of M and X.

Ans. O.N of MX_3 = 0

O.N of M = +3

O.N of X_3 = -3

iii. Why the oxidation number of oxygen in OF_2 is +2?

Ans. The oxidation number of oxygen in normal oxide is -2. But in OF_2 is +2, because "F" is more electronegative atom than oxygen hence oxidation number of two "F" atoms is "-2" and that of oxygen is +2.

iv. In H_2S, SO_2 and H_2SO_4 the sulphur atom has different oxidation number. Find out the oxidation number of sulphur in each compound.

Ans. H_2S

$(+1)2+(S)1 = 0$

$S + 2 = 0$

O.N of S = -2

SO_2

$(S)1+(-2)2 = 0$

$S - 4 = 0$

O.N of S = +4

H_2SO_4

$(+1)2 +(S) 1 + (-2)4 = 0$

$+2 + S - 8 = 0$

$S - 6 = 0$

O.N of S = +6

v. An element X has oxidation state 0. What will be its oxidation state when it gains three electrons?

Ans. -3 will be the oxidation state of × when it gais three electrons.

vi. An element in oxidation state + 7 gains electrons to be reduced to oxidation state +2. How many electrons did it accept?

Ans. Five electrons were accepted by it.

vii. If the oxidation state of an element changes from +5 to -3. Has it been reduced or oxidized? How many electrons are involved in this process?

Ans. It has been reduced. This element gains 8 electrons.

7.3 OXIDIZING AND REDUCING AGENTS

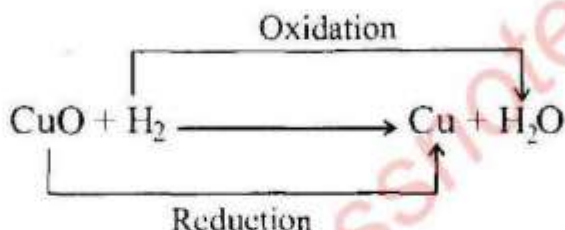
7.4 OXIDATION - REDUCTION REACTIONS

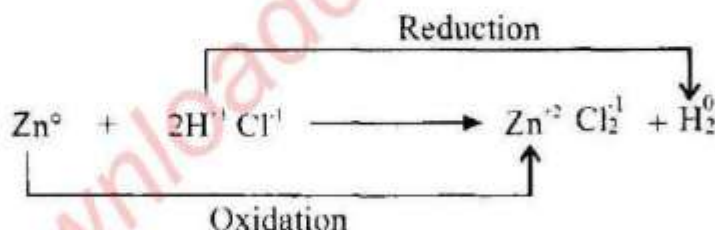
Q.3.(a) What are redox reaction? Explain with examples.

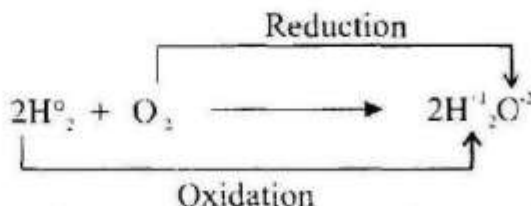
(b) Define the oxidation and reduction in terms of oxidation states.

(c) Define the oxidizing and reducing agents with examples.

Ans.(a) Redox [Oxidation - Reduction] Reactions: The reactions in which oxidation as well as reduction take place are called redox reactions.

Examples No. 1:

Explanation: In the above reaction "CuO" changes to "Cu" by losing oxygen hence it is reduction while "H_2" gains oxygen to give water, hence it is oxidation.

Example No. 2:

Explanation: In above reaction, the oxidation state of "Zn" increase from "0" to +2 hence it is oxidation while oxidation state of H changes from "+" to "0" hence it is reduction.

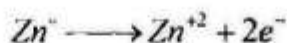
Example No. 3:

Explanation: In above reaction, the oxidation state of hydrogen increase from "0" to "+" hence it is oxidation while the oxidation state of oxygen decreases from 0 to -2, hence, it is reduction.

Oxidation is 'losing electrons in a chemical reaction'
Reduction is 'gaining electrons in a chemical reaction'
Reducing agent - is a substance that oxidizes itself and reduces other.
Oxidizing agent - is a substance that reduces itself and oxidizes other.

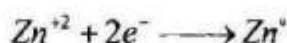
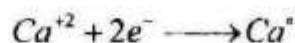
(b) The oxidation and reduction in terms of oxidation states:

Oxidation: The increase in oxidation state is called oxidation e.g.

$$Zn^{\circ} \longrightarrow Zn^{+2} + 2e^-$$

$$Al^{\circ} \longrightarrow Al^{+3} + 3e^-$$

Reduction: The decrease in the oxidation state is called reduction. e.g.

$$Ca^{+2} + 2e^- \longrightarrow Ca^{\circ}$$

$$Zn^{+2} + 2e^- \longrightarrow Zn^{\circ}$$

(c) The oxidizing and reducing agents:

Oxidizing agent [Oxidants]: The substances which help the oxidation to occur are called oxidizing agents or oxidants.

They oxidise the other substances by taking electrons from them and themselves get reduced by gaining the electrons.

Examples:

1. Mostly non-metals are oxidizing agents.
2. Acidified $KMnO_4$ and $K_2Cr_2O_7$ are oxidizing agents.

Reducing Agents [Reductants]: The substances which help the reduction to occur are called reducing agents.

They reduce the other substances by giving electrons to them and themselves get oxidised.

Examples:

1. Almost all metal are good reducing agents.
2. CO, SO_2, H_2 KI are also reducing agents.

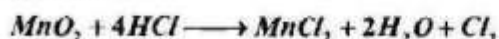
Test yourself 7.3:

i. In the following reaction, how can you justify that H_2S is oxidized and SO_2 is reduced.

$$SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$$

Ans. SO_2 is oxidised from SO_2 to S. due to removal of oxygen.
"H_2S" is oxidised to H_2O due to addition of oxygen

ii. The reaction between MnO_2 and HCl is a redox reaction written as balance chemical equation.

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

Find out: a. The substance oxidized b. The substance reduced
c. The substance which acts as oxidizing agent
d. The substance which acts as reducing agent

Ans. $Mn^{+4}O_2^{-2} + 4H^{+1}Cl^{-1} \longrightarrow Mn^{+2}Cl_2^{-1} + 2H_2^{+1}O^{-2} + Cl_2^0$

(a) Cl^{-1} is oxidised to Cl_2^0
(b) M_n^{+4} is reduced to Mn^{+2}
(c) Oxidising agent, MnO_2

(d) Reducing agent HCl

iii. **The following reactions are redox reactions.**
Find out the element which has been reduced and the element which has been oxidized.

a. $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

b. $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$

c. $H_2S + Cl_2 \longrightarrow 2HCl + S$

Ans. (a) $Zn^0 + Cu^{+2}SO_4^{-2} \longrightarrow Zn^{+2}SO_4^{-2} + Cu^0_{(s)}$

Zinc oxidises from Zn^0 to Zn^{+2}. Copper reduces from Cu^{+2} to Cu^0

(b) $Cu^0 + 2Ag^{+1}NO_3^{-1} \longrightarrow Cu^{+2}(NO_3^{-1})_2 + 2Ag^0$

Copper oxidised from Cu^0 to Cu^{+2} and silver reduced from Ag^{+1} to Ag^0

(c) $H_2^{+1}S^{-2} + Cl_2^0 \longrightarrow 2H^{+1}Cl^{-1} + S^0$

Sulphur oxidises from S^{-2} to S^0 and chlorine reduces from Cl_2^0 to Cl^-

iv. **Why the following reaction is not a redox reaction. Explain with reasons?**

$NaOH + HCl \longrightarrow NaCl + H_2O$

Ans. $Na^{+1}O^{-2}H^{+1} + H^{+1}Cl^{-1} \longrightarrow Na^{+1}Cl^{-1} + H_2^{+1}O^{-2}$

This reaction is not a redox reaction because no oxidation or reduction take place.

7.5 ELECTROCHEMICAL CELLS

Q.4. What is meant by electrolytes? Describe their types with example?

Ans. Electrolytes: The substances, which can conduct electricity in their solutions or molten states, are called electrolytes. For example, solution of salts, acids or bases are good electrolytes. The electricity cannot pass through solid NaCl but its equeous solution or molten NaCl are good electrolytes. Electrolytes are classified into two groups depending upon their extent of ionization in solution.

Types of Electrolytes:

1. **Strong Electrolytes:** The electrolytes which ionize completely in solution and produce more ions are called strong electrolytes. Example of strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4 etc.

$$NaOH_{(s)} + H_2O \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

2. **Weak Electrolytes:** The substances which ionize to a small extent when dissolved in water and could not produce more ions are called weak electrolytes. Acetic acid (CH_3COOH) and $Ca(OH)_2$ when dissolved in water, ionize to a small extent: are good examples of weak electrolytes. Weak electrolytes do not ionize completely. For example, ionization of acetic acid in water produces less ions.

$$CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

As a result the weak electrolyte is a poor conductor of electricity.

3. **Non Electrolytes:** The substances, which do not ionize in solution and do not

allow the current to pass through their solutions, are called non-electrolytes. For example, sugar solution and benzene are non-electrolytes.

Q.5. Discuss the construction and working of electrolytic cell.

Ans. Electrolytic cells: The type of electrochemical cell in which a non-spontaneous chemical reaction takes place when electric current is passed through the solution is called an electrolytic cell.

Examples:

Examples of these cells are Down cells, Nelson's cell and electrolysis of water.

Construction of an electrolytic cell:

An electrolytic cell consists of a solution of an electrolyte, two electrodes (anode and cathode) that are dipped in the solution and connected to the battery. The electrode connected to positive terminal is called anode and electrode connected to the negative terminal is called cathode as shown in figure.

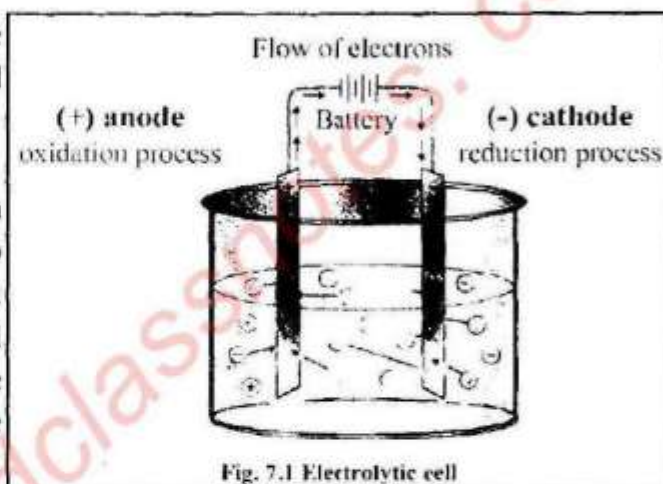
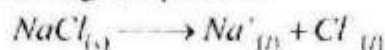



Fig. 7.1 Electrolytic cell

Working of an Electrolytic cell: When electric current is applied from battery. The ions in the solution migrate to their respective electrodes. The anions, which are negatively charged, move towards the anode and discharge there by losing their electrons. Thus oxidation takes place at anode. While cations, which are positively charge ions, move towards cathode. Cations gain electrons from the electrode and as a result reduction takes place at cathode. For example, when fused salt of sodium chloride is electrolysed the following reactions take place during this process.

$$NaCl_{(s)} \longrightarrow Na^{+}_{(l)} + Cl^{-}_{(l)}$$

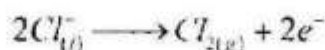
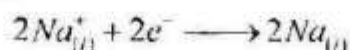
Oxidation reaction at anode:

$$2Cl^{-}_{(l)} \longrightarrow Cl_{2(g)} + 2e^{-}$$

Reduction reaction at cathode:

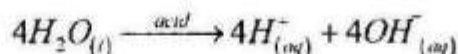
$$2Na^{+}_{(l)} + 2e^{-} \longrightarrow 2Na_{(l)}$$

Overall reaction:

$$2Na^{+}_{(l)} + 2Cl^{-}_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2(g)}$$

Q.6. Describe the electrolysis of water.

Ans. Electrolysis of water: Pure water is a very weak electrolyte. It ionizes to a very small extent. The concentrations of hydrogen ions (H^+) and hydroxyl ions (OH^-) are both

10^{-7} mol dm^{-3} respectively. When a few drops of an acid are added in water, its conductivity improves.

$$4H_2O_{(l)} \xrightarrow{acid} 4H^+_{(aq)} + 4OH^-_{(aq)}$$

When an electric current is passed through this acidified water, OH^- anions move towards positive electrode (anode) and H^+ cations move towards negative electrodes (cathode) and discharge takes place at these electrodes. They produce oxygen and hydrogen gases respectively at anode and cathode as shown in figure.

The redox reaction taking place in the electrolytic bath can be shown as following.

Electrolytic cell showing electrolysis of water

Oxidation reaction at anode:

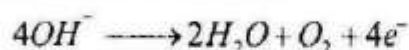
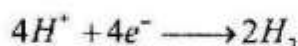
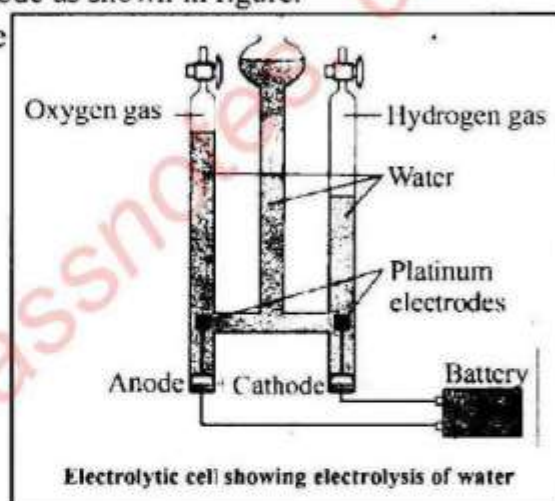
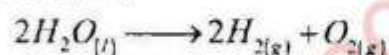
$$4OH^- \longrightarrow 2H_2O + O_2 + 4e^-$$

Reduction reaction at cathode:

$$4H^+ + 4e^- \longrightarrow 2H_2$$

$$4H \longrightarrow 2H_2$$

Overall reaction:

$$2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}$$

Q.7. What is galvanic or voltaic cell? Describe the construction of Daniel cell.

Ans. Galvanic cell OR Voltaic cell:

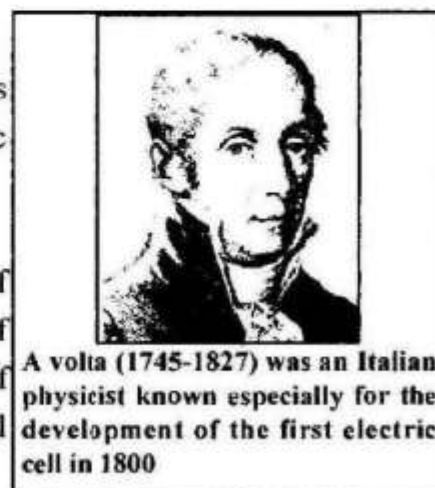
The electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called galvanic or voltaic cell.

Construction of a Daniel cell:

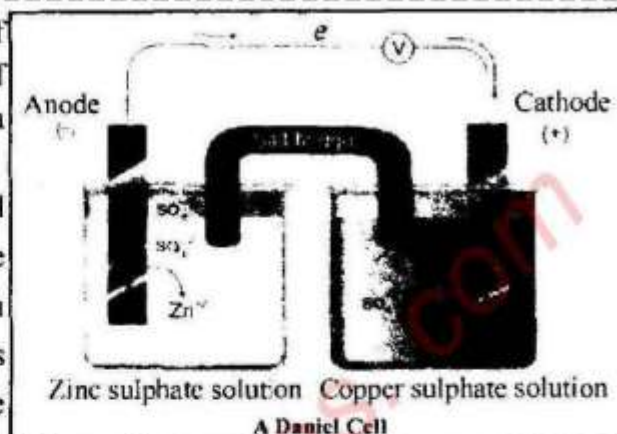
A volta (1745-1827) was an Italian physicist known especially for the development of the first electric cell in 1800

A galvanic cell consists of two cells, each called as half cell, connected electrically by a salt-bridge. In each of the half cell, an electrode is dipped in 1M solution of its own salt and connected by a wire to an external circuit, as shown in figure shows a typical galvanic cell.

The left half cell consists of an electrode of zinc metal dipped in 1M solution of zinic sulphate.

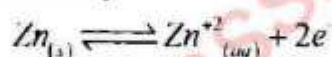
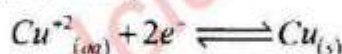
The right half cell is a copper electrode dipped in 1M solution of copper sulphate. Salt bridge is a U shaped glass tube. It consists of saturated solution of strong electrolyte supported in a jelly type material. The ends of the U tube are sealed with a porous

material like glass wool. The function of the salt bridge is to keep the solution of two half cells neutral by providing a pathway for migration of ions.

A Daniel Cell

Working of Daniell cell: The Zn metal has tendency to lose electrons more readily than copper. As a result oxidation takes place at Zn electrode. The electrons flow from Zn electrode through the external wire in a circuit to copper electrode. These electrons are gained by the copper ions of the solution and copper atoms deposit at the electrode. The respective oxidation and reduction processes going on at two electrodes are as follows.

Half cell reaction at anode (oxidation):

$$Zn_{(s)} \rightleftharpoons Zn^{+2}_{(aq)} + 2e^-$$

Half cell reaction at cathode (reduction):

$$Cu^{+2}_{(aq)} + 2e^- \rightleftharpoons Cu_{(s)}$$

Overall galvanic reaction is the sum of these two half-cell reactions

$$Zn_{(s)} + Cu^{+2}_{(aq)} \rightleftharpoons Zn^{+2}_{(aq)} + Cu_{(s)}$$

As a result of redox reaction electric current is produced, which is used for starting automobiles, running calculators and toys and to lit the bulbs.

Q.8. Differentiate the electrolytic and galvanic cells.

Ans. A comparison of electrolytic and Galvanic cells:

Electrolytic Cell	Galvanic Cell
1. It consists of one complete cell, connected to a battery.	1. It consist of two half cells connected through a salt bridge.
2. Anode has positive charge while cathode has negative charge.	2. Anode has negative charge while cathode has positive charge.
3. Electrical energy is converted into chemical energy.	3. Chemical energy is converted into electrical energy.
4. Current is used for a non spontaneous chemical reaction to take place.	4. Redox reaction takes place spontaneously and produces electric current.

Test yourself 7.4:

i. **Why are the strong electrolytes termed as good conductors?**

Ans. Strong electrolytes are fully ionized in aqueous solution and thus conduct electric current to a

large extent, hence are called good conductors.

ii. Does non-electrolytes form ions in solution?

Ans. NO, non-electrolytes do not form ions in solution.

iii. What is difference between a strong electrolyte and a weak electrolyte?

Ans. **Strong electrolytes:** The compounds which are fully ionized in aqueous solution and conduct electric current to a large extent e.g. aqueous solution of NaCl.

Weak electrolytes: The compounds which are partially ionized in aqueous and thus conduct electric current to a small extent are called weak electrolytes e.g. aqueous solution of CH_3COOH.

iv. Identify a strong or weak electrolyte among the following compounds:
$CuSO_4$, H_2CO_3, $Ca(OH)_2$, HCl, $AgNO_3$

Ans.

Strong electrolytes	Weak electrolytes
$CuSO_4$, HCl $AgNO_3$	H_2CO_3, $Ca(OH)_2$

v. Which force drives the non-spontaneous reaction to take place?

Ans. Non-spontaneous reaction takes place when electric current is passed through the solution.

vi. Which type of chemical reaction takes place in electrolytic cell?

Ans. Non-spontaneous chemical reaction takes place in electrolytic cell.

vii. What type of reaction takes place at anode in electrolytic cell?

Ans. Oxidation reaction takes place at anode.

viii. Why the positively charged electrode is called anode in electrolytic cell?

Ans. The electrode which is connected to the positive terminal of battery is called anode. The oxidation reaction takes place at this electrode.

ix. In the electrolysis of water, towards which terminal H^+ ions move?

Ans. In the electrolysis of water H^+ ions move towards cathode.

x. In the electrolysis of water, where is the oxygen produced?

Ans. In the electrolysis oxygen is produced at anode.

xi. Towards which electrode of the electrolytic cell moves the cations and what does they do there?

Ans. The cations move towards the cathode and gains electrons at cathode to become neutral.

xii. How the half cells of a galvanic cell are connected? What is function of salt bridge?

Ans. Half cells are connected by salt bridge. The function of the salt bridge is to keep the solutions of two half calls neutral. By providing a pathway for migration of ions.

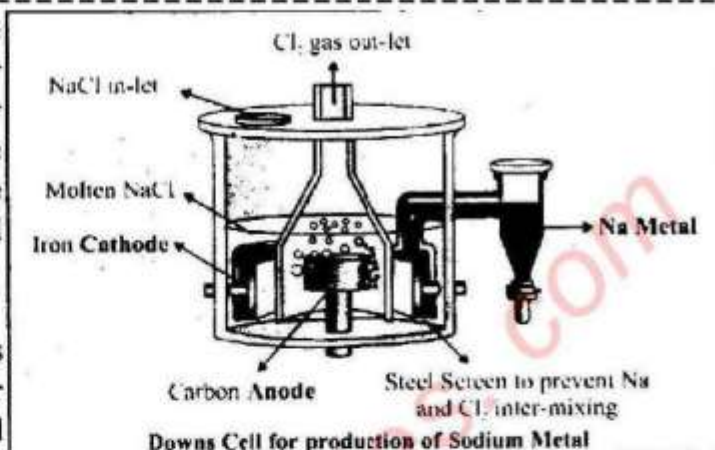
7.6 ELECTROCHEMICAL INDUSTRIES

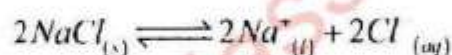
Q.9. Discuss the manufacture of sodium metal from fused NaCl by using Downs cell. OR

Discuss the electrolysis of fused NaCl by using Downs cell.

Ans. Manufacture of sodium metal from fused NaCl: On the industrial scale molten sodium metal is obtained by the electrolysis of fused NaCl in the Downs cell. This electrolytic cell is a circular furnace. In the center there is a large block of graphite, which acts as an anode while cathode around it is made of iron as shown in figure.

Working of Downs cell: The fused NaCl produces Na^+ and Cl^- ions, which migrate to their respective electrodes on the passage of electric current. The electrodes are separated by steel gauze to prevent the contact between the products. The Cl^- ions are oxidized to give Cl_2 gas at the anode. It is collected over the anode within an inverted

Downs Cell for production of Sodium Metal

cone-shaped structure. While Na^+ are reduced at cathode and molten Na metal floats on the denser molten salt mixture from where it is collected in a side tube. Following reactions take place during the electrolysis of the molten sodium chloride.

Molten NaCl ionizes as:

$$2NaCl_{(s)} \rightleftharpoons 2Na^+_{(l)} + 2Cl^-_{(aq)}$$

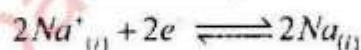
Half-cell reaction at anode (oxidation)

$$2Cl^-_{(g)} \rightleftharpoons Cl_{2(g)} + 2e^-$$

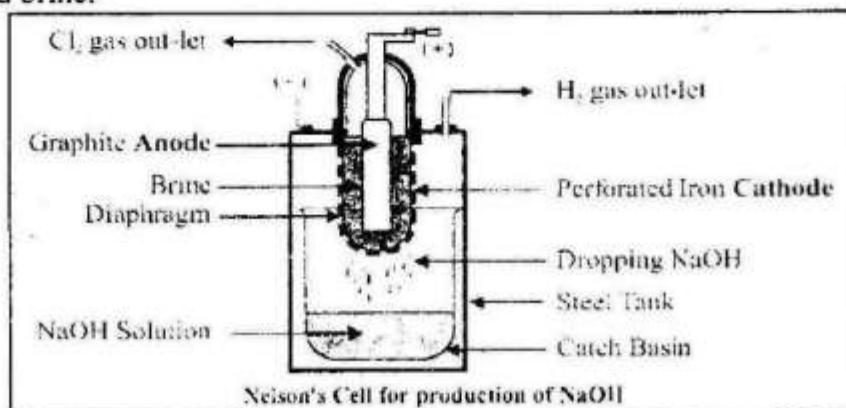
Half-cell reaction at cathode (reduction)

$$2Na^+_{(l)} + 2e^- \rightleftharpoons 2Na_{(l)}$$

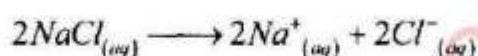
Overall glavanic reaction is the sum of these two half-cell reactions

$$2NaCl_{(fused)} \rightleftharpoons Cl_{2(g)} + 2Na_{(l)}$$

Q.10. Discuss commercial preparation of NaOH from Brine by using Nelson cell.

Ans. Manufacture of NaOH from Brine: On industrial scale caustic soda, sodium hydroxide NaOH, is produced in Nelson's cell by the electrolysis of equeous solution of NaCl called brine.

Nelson's Cell for production of NaOH

Construction of Nelson cell: It consists of a steel tank in which graphite anode is suspended in the center of a U shaped perforated iron cathode. This iron cathode is internally lined with asbestos diaphragm. Electrolyte brine is present inside the iron cathode.

Working of Nelson's cell: Aqueous solution of sodium chloride consists of Na^+, Cl^-, H^+ and OH^- ions. These ions move towards their respective electrodes and redox reactions take place at these electrodes. When electrolysis takes place Cl^- ions are discharged at anode and Cl_2 gas rises into the dome at the top of the cell. The H^+ ions are discharged at cathode and H_2 gas escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.

Ionization of Brine:

$$2NaCl_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2Cl^-_{(aq)}$$

Ionization of water:

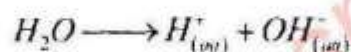
$$H_2O \longrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

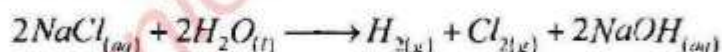
Reaction at anode (oxidation):

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$$

Reaction at cathode (reduction):

$$4OH^- \longrightarrow 2H_2O + O_2$$

$$2Na^+ + 2OH^- \longrightarrow 2NaOH$$

Overall cell reaction of this process:

$$2NaCl_{(aq)} + 2H_2O_{(l)} \longrightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$$

Test yourself 7.5:

i. **Anode of Downs cell is made of a non-metal, what is its name? What is the function of this anode?**

Ans. The anode of down cell is composed of graphite. The oxidation of chloride ions takes place at anode.

ii. **Where does the sodium metal is collected in Downs cell?**

Ans. Sodium is produced at iron cathode and is collected in a side tube.

iii. **What is the name of the by -product produced in the Downs cell?**

Ans. Chlorine is produced as by product in Down's cell.

iv. **Are anodes of Downs cell and Nelson cell made of same element? If yes, what is its name?**

Ans. The anodes of Down's cell and Nelson cell are made up of graphite.

v. **What is the shape of cathode in Nelson's cell? Why is it perforated?**

Ans. Perforated iron cathode. It is perforated because sodium hydroxide solution has to percolates into a cath basin.

vi. **Which ions are discharged at cathode in Nelson's cell and what is produced at cathode?**

Ans. Hydrogen ions are discharged at cathode in Nelson's cell. It is perforated.

7.7 CORROSION AND ITS PREVENTION

Q.11. What is meant by rusting or corrosion? Discuss the rusting of iron.

Ans. Corrosion:

It is slow and continuous eating away of a metal by the surrounding medium. It is a redox chemical reaction that takes place by the action of air and moisture with the metals.

Example:

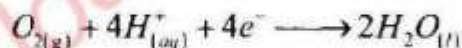
The most common example of corrosion is rusting of iron.

Rusting of iron: Corrosion is a general term but corrosion of iron is called rusting. The important condition for rusting is moist air (air having water vapours in it). There will be no rusting in water vapours free of air or air free of water.

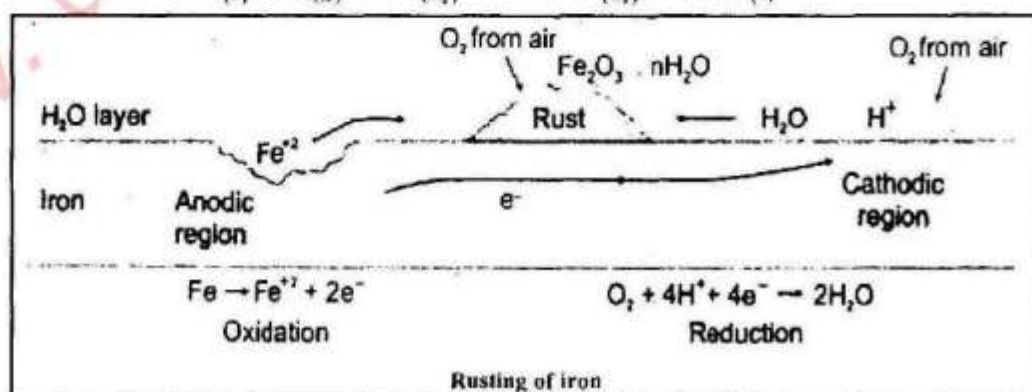
Chemistry of iron: Stains and dents on the surface of the iron provide the sites for this process to occur. This region is called anodic region and following oxidation reaction takes place.

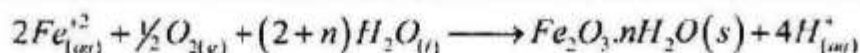
$$2Fe \longrightarrow 2Fe^{+2} + 4e^-$$

The loss of electrons damage the object. The free electrons move through iron sheet, until they reach to a region of relatively high O_2 concentration near the surface surrounded by water layer as shown in figure. This region acts as cathode and electrons reduce the oxygen molecule in the presence of H^+ ions:

$$O_{2(g)} + 4H^+_{(aq)} + 4e^- \longrightarrow 2H_2O_{(l)}$$

The H^+ ions are provided by the carbonic acid, which is formed because of presence of CO_2 in water. That is why acidic medium accelerates the process of rusting. The overall redox process is completed without the formation of rust.

$$2Fe_{(s)} + O_{2(g)} + 4H^+_{(aq)} \longrightarrow 2Fe^{+2}_{(aq)} + 2H_2O_{(l)}$$

Rusting of iron

The Fe^{+2} formed spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3 . nH_2O$ which is called rust. It is also redox reaction.

$$2Fe^{+2}_{(aq)} + \frac{1}{2}O_{2(g)} + (2+n)H_2O_{(l)} \longrightarrow Fe_2O_3.nH_2O(s) + 4H^+_{(aq)}$$

The rust layer of iron is porous and does not prevent further corrosion. Thus rusting continues until whole the piece of irons is eaten up

Technology

Does Aluminum Rust? Aluminium corrodes but it does not rust. Rust refers only to iron and steel corrosion. A very hard material aluminium oxide protects the aluminium from further corrosion. In comparison to that when iron corrodes, its color changes and produces large red flakes known as rust. Unlike aluminium oxide, the expanding and flaking of rust exposes new metal surface to further rusting.

Q.12. How the process of corrosion or rusting can be prevented. Explain it:

Ans. Prevention of Corrosion:

Removal of stains: The regions of stains in an iron rod act as the site for corrosion. If the surface of iron is properly cleaned and stains are removed, it would prevent corrosion.

Paints and greasing: Polishing or painting of the surface can prevent the rusting of iron. With development of technologies, modern paints contain a combination of chemicals called stabilizers that provide protection against the corrosion in addition to prevention against the weathering and other atmospheric effects.

Alloying: Alloy is a homogeneous mixture of one metal with one or more other metals or non-metals. Alloying of iron with other metals has proved to be very successful technique against rusting. The best example of alloying is the 'stainless steel', which is a good combination of iron, chromium and nickel.

Metallic coating: The best method for protection against the corrosion of metals exposed to acidic conditions is coating the metals with other metals. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion. It is the most widely applied technique in the food industry where food is "tin-packed". The containers of iron are coated with tin or chromium to give it a longer life. Metallic coating can take place by physical as well as electrolytic methods.

Physical Methods (galvanizing and tin coating):

(a) Zinc coating or Galvanizing: The process of coating a thin layer of zinc on iron is called galvanizing. This process is carried out by dipping a clean iron sheet in a zinc chloride bath and then heating it. After this iron sheet is removed, rolled into molten zinc metal bath and finally air-cooled. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

Tin Coating: It involves the dipping of the clean sheet of iron in a bath of molten tin and then passing it through hot rollers. Such sheets are used in the beverage and food cans. The tin protects the iron only as long as its protective layer remains intact. Once it is broken and the iron is exposed to the air and water, a galvanic cell is established and iron

rusts rapidly.

Test yourself 7.6:

i. What is the difference between corrosion and rusting?

Ans. Corrosion is slow and continuous eating away of a metal by the surrounding medium. It is a redox chemical reaction that takes place by the action of air and moisture with metal. The corrosion is a general term but corrosion of iron is called rusting.

ii. What happens to iron in the rusting process?

Ans. Iron is converted into hydrated iron oxide. $Fe_2O_3 . nH_2O$.

iii. Rusting completes in how many redox reactions?

Ans. Rusting process is complete in two redox reactions.

iv. Explain the role of O_2 in rusting?

Ans. Oxygen converts iron into iron oxide. (Fe_2O_3).

v. State the best method for protection of metal from corrosion.

Ans. Zinc coating or galvanizing is the best method for protection of metal from corrosion.

vi. What do you mean by galvanizing?

Ans. **Galvanizing:** The process of creating a thin layer of zinc on iron is called galvanizing.

vii. What is the advantage of galvanizing?

Ans. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

viii. Why tin plated iron is rusted rapidly when tin layer is broken?

Ans. Once tin layer is broken the iron is exposed to the air and water, and iron rusts rapidly.

ix. Name the metal which is used for galvanizing iron?

Ans. Zinc metal is used for galvanizing iron.

Q.13. What is electroplating? Write down procedure of electroplating.

Ans. Electroplating: It is depositing of one metal over the other by means of electrolysis.

Importance: This process is used to protect metals against corrosion and to improve their appearance.

Principle: Principle of electroplating is to establish an electrolytic cell in which anode is made of the metal to be deposited and cathode of the object on which metal is to deposit. The electrolyte is in equeous solution of a salt of the respective metal.

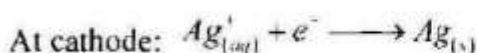
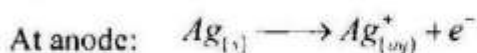
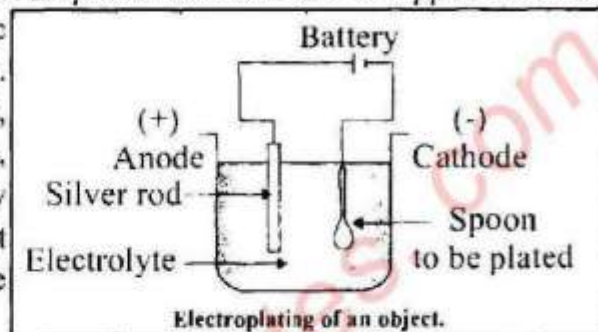
Procedure for Electroplating: In this process, the object to be electroplated is cleaned with sand, washed with caustic soda solution and finally it is thoroughly washed with water. The anode is made of the metal, which is to be deposited like Cr, Ni. The cathode is made up of the object that is to be electroplated like some sheet made up of iron. The electrolyte in this system is a salt of the metal being deposited. The electrolytic tank is made of cement, glass or wood in which anode and cathode are suspended. These electrodes are connected with a battery. When the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and discharge or deposit on the cathode (object). As a result of this discharge, a thin layer of metal deposits on the object, which is then pulled out and cleaned.

Q.14. Discuss the electroplating of silver.

Ans. Electroplating of Silver: The electroplating of silver is carried out by establishing an electrolytic cell. The pure piece of silver strip acts as anode that is dipped in silver nitrate solution. The cathode is the metallic object to be coated such as silver spoon. When the current is passed through the cell, the anode dissolves to produce Ag^+ ions, that migrate towards the cathode where they are discharged and deposited on the object e.g. spoon. The chemical reaction can be represented as:

Electroplating of an object.

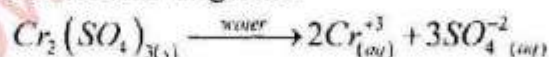
At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + e^-$

At cathode: $Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}$

Common examples of silver plating are table wares, cutlery, jewelry and steel objects.

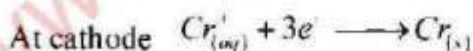
Q.15. Explain the electroplating of chromium.

Ans. Electroplating of Chromium: The electroplating of chromium is carried out in the same way as that of silver. The object to be electroplated is dipped in aqueous solution of chromium sulphate containing a little sulphuric acid, that acts as an electrolyte. The objects to be electroplated acts as cathode while anode is made of antimonial-lead. The electrolyte ionizes and provides Cr^{3+} ions, which reduce and deposit at cathode.

Electrolyte prdouces the following ions:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{water} 2Cr^{+3}_{(aq)} + 3SO^{-2}_{4\,(aq)}$$

Reactions at the electrodes are as follows:

At anode $4OH_{(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$

At cathode $Cr^+_{(aq)} + 3e \longrightarrow Cr_{(s)}$

For practical convenience, the steel is usually plated first with nickel or copper and then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off. The nickel or copper provides adhesion and then chromium deposited over the adhesive layer of copper lasts longer. This type of electroplating resists corrosion and gives a bright silvery appearance to the object.

Q.16. Describe the following.

(a) Electroplating of zinc *(b) Electroplating of tin*

(c) ***Electrolytic refining of copper***

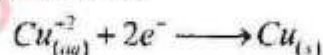
Ans.(a) Electroplating of zinc: The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales. Then, the zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied, which results in depositing of zinc on the target metal i.e. cathode.

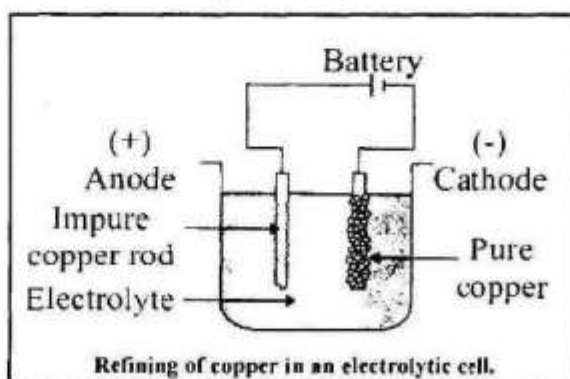
(b) Electroplating of tin: Tin is usually electroplated on steel by placing the steel into a container containing a solution of tin salt. The steel is connected to an electrical circuit, acting as cathode. While the other electrode made of tin metal acts as anode. When an electrical current passes through the circuit, tin metal ions present in the solution deposit on steel.

(c) Electrolytic refining of copper: Impure copper is refined by the electrolytic method in the electrolytic cell. Impure copper acts as anode and a pure copper plate acts as cathode as shown in figure. Copper sulphate solution in water is used as an electrolyte. Oxidation reaction takes place at the anode. Copper atoms from the impure copper lose electrons to the anode and dissolve in solution as copper ions.

$$Cu_{(s)} \longrightarrow Cu^{+2}_{(aq)} + 2e^-$$

Reduction reaction takes place at the cathode. The copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.

$$Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

In the process impure copper is eaten up and purified copper atoms deposit on the cathode.

Refining of copper in an electrolytic cell.

Test yourself 7.7:

i. **Define electroplating?**

Ans. **Electroplating**: The process of depositing a metal at the surface of another metal by the help of electroplating.

ii. **How electroplating of zinc is carried out?**

Ans. The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid,

in order to remove any rust of surface scales. Next, the zinc is deposited on the metal by immersing it in chemical bath containing electrolyte zinc sulphate. A current is applied, which results in zinc being deposited on the target metal.

iii. Which material is used to make cathode in electroplating?

Ans. The cathode is made up of object.

iv. Why is the anode made up of a metal to be deposited during electrolysis?

Ans. Because metal (anode) dissolves in solution to give metal ions which are deposited at cathode to make layer.

Key Points

- Oxidation is addition of oxygen or removal of hydrogen or loss of electrons by an element and as a result oxidation number increases.
- Reduction is addition of hydrogen or removal of oxygen or gain of electrons by an element and as a result oxidation number decreases.
- Oxidation number is the apparent charge on an atom. It may be positive or negative.
- Oxidizing agents are the species that oxidize the other element and reduce themselves. Non-metals are oxidizing agents.
- Reducing agents are species that reduce the other elements and oxidize themselves. Metals are reducing agents.
- Chemical reactions in which the oxidation state of species change are termed as redox reaction. A redox reaction involves oxidation and reduction processes taking place simultaneously.
- Redox reactions either take place spontaneously and produce energy or electricity is used to drive the reaction.
- The process in which electricity is used for the decomposition of a chemical compound is called electrolysis. It takes place in electrolytic cells such as Downs cell and Nelson's cell.
- Galvanic cells are those in which spontaneous reactions take place and generate electric current. They are also called voltaic cells.
- Sodium metal is manufactured from fused sodium chloride in the Downs cell.
- NaOH is manufactured from brine in Nelson's cell.
- Corrosion is slow and continuous eating away of a metal by the surrounding medium. The most common example of corrosion is rusting of iron.
- The rusting principle is electrochemical redox reaction, in which iron behaves as anode. Iron is oxidized to form rust $Fe_2O_3 . nH_2O$.
- Corrosion can be prevented by many methods. The most important is electroplating.
- Electroplating is depositing of one metal over the other by means of electrolysis .
- Iron can be electroplated by tin, zinc, silver or chromium.

Exercise (Solved)

☆ Multiple Choice Questions

Put a (✓) on the correct answer.

1. Spontaneous chemical reactions take place in:
(a) Electrolytic cell (b) Galvanic cell (c) Nelson's cell (d) Downs cell

2. Formation of water from hydrogen and oxygen is:
(a) Redox reaction (b) Acid-base reaction
(c) Neutralization (d) Decomposition

3. Which one of the following is not an electrolytic cell?
(a) Downs cell (b) Galvanic cell (c) Nelson's cell (d) Both a and c

4. The oxidation number of chromium in $K_2Cr_2O_7$ is:
(a) +2 (b) +6 (c) +7 (d) +14

5. Which one of the following is not an electrolyte?
(a) Sugar solution (b) Sulphuric acid solution
(c) Lime solution (d) Sodium chloride solution

6. The most common example of corrosion is:
(a) Chemical decay (b) Rusting of iron
(c) Rusting of aluminium (d) Rusting of tin

7. Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:
(a) Cl_2 (b) H_2 (c) O_3 (d) O_2

8. During the formation of water from hydrogen and oxygen, which of the following does not occur:
(a) Hydrogen has oxidized (b) Oxygen has reduced
(c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent

9. The formula of rust is:
(a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3 (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$

10. In the redox reaction between Zn and HCl, the oxidizing agent is:
(a) Zn (b) H^+ (c) Cl^- (d) H_2

Answers: **1.** Galvanic cell **2.** Redox reaction **3.** Galvanic cell **4.** +6
5. Sugar solution **6.** Rusting of iron **7.** H_2
8. Hydrogen behaves as oxidizing agent **9.** $Fe_2O_3.nH_2O$ **10.** H^+

☆ Short Answer Questions.

1. Define oxidation in terms of electrons. Give an example.

Ans. For answer see Q. 1

2. Define reduction in terms of loss or gain of oxygen or hydrogen. Give an example.

Ans. For answer see Q. 1

3. What is the difference between valency and oxidation state?

Ans. Valency is combining capacity of an element with other elements while oxidation state or oxidation number (O.N.) is the apparent charge assigned to an atom of an element in a molecule or in an ion.

4. Differentiate between oxidizing and reducing agents

Ans. For answer see Q. 3

5. Differentiate between strong and weak electrolytes.

Ans. For answer see Q. 4

6. How is electroplating of tin on steel carried out?

Ans. For answer see (electroplating of tin) Q. 16

7. Why is steel plated with nickel before the electroplating of chromium.

Ans. For answer see Q. 15

8. How can you explain, that following reaction is oxidation in terms of increase of oxidation number? $Al^{0} \longrightarrow Al^{+3} + 3e^{-}$

Ans. Oxidation state of "Al" increases from "0" to +3.

9. How can you prove with an example that conversion of anion to an atom is an oxidation process?

Ans. The increase in oxidation state is called oxidation. e.g. $Al^{0} \longrightarrow Al^{+3} + 3e^{-}$

10. Why does the anode carry negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

Ans. For answer see Q. 5, Q. 7

11. Where do the electrons flow from Zn electrode in Daniel's cell?

Ans. For answer see Q. 7

12. Why do electrodes get their names 'anode' and cathode in galvanic cell?

Ans. For answer see Q. 7

13. What happens at the cathode in a galvanic cell?

Ans. For answer see Q. 7

14. Which solution is used as an electrolyte in Nelson's cell?

Ans. For answer see Q. 10

15. Name the by-products produced in Nelson's cell?

Ans. For answer see Q. 10

16. Why is galvanizing done?

Ans. For answer see Q. 12

17. Why is an iron grill painted frequently?

Ans. Inorder to prevent from rusting.

18. Why is O_2 necessary for rusting?

Ans. For answer see Q. 12

19. In electroplating of chromium, which salt is used as an electrolyte?

Ans. For answer see Q. 15

20. Write the redox reaction taking place during the electroplating of chromium?

Ans. For answer see Q. 15

21. In electroplating of silver, from where do Ag^+ ions come and where do they deposit?

Ans. For answer see Q. 14

22. What is the nature of electrode used in electroplating of chromium?

Ans. For answer see Q. 15

Long Answer Questions

1. Describe the rules for assigning the oxidation state

Ans. For answer see Q. 2

2. Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2\underline{S}O_4$ (b) $Ag\underline{N}O_3$ (c) $K\underline{Mn}O_4$

(d) $K_2\underline{Cr}_2O_7$ (e) $H\underline{N}O_2$

Ans. For answer see Q. 2

3. How can a non-spontaneous reaction be carried out in an electrolytic cell. Discuss in detail.

Ans. For answer see Q. 5

4. Discuss the electrolysis of water.

Ans. For answer see Q. 6

5. Discuss the construction and working of a cell in which electricity is produced.

Ans. For answer see Q. 7

6. How can we prepare NaOH on commercial scale. Discuss its chemistry along with the diagram.

Ans. For answer see Q. 10

7. Discuss the redox reaction taking place in the rusting of iron in detail.

Ans. For answer see Q. 11

8. Discuss, why galvanizing is considered better than tin plating.

Ans. For answer see Q. 12

9. What is electroplating? Write down procedure of electroplating.

Ans. For answer see Q. 13

10. What is the principle of electroplating? How is electroplating of chromium carried out?

Ans. For answer see Q. 13, Q.15

OBJECTIVE TYPE QUESTIONS (MCQ's+SHORT ANSWER) FROM PREVIOUS ANNUAL PAPERS OF ALL SECONDARY BOARDS (LAHORE, GUJRANWALA, FAISALABAD, MULTAN, SAHIWAL, SARGODHA, RAWALPINDI, D.G. KHAN AND BAHAWALPUR)

7.1 Oxidation and Reduction Reactions

7.2 Oxidation State and Rules for Assigning Oxidation State

7.3 Oxidizing and Reducing Agents

☆ **Tick the correct answer.**

1. **Addition of oxygen during chemical reaction is called:** (LHR. GI, & GII)
   (A) Reduction (B) Oxidation (C) Evaporation (D) Conduction
2. **In _____ branch of chemistry, the relationship between electricity and chemical reaction is studied:** (MLN. GII)
   (A) Organic Chemistry (B) In Organic Chemistry
   (C) Electrochemistry (D) Industrial Chemistry
3. **The equation $2H^+ + 2e^- \longrightarrow H_2$ is indicating the process of:** (MLN. GII)
   (A) Oxidation (B) Reduction (C) Redox (D) Decomposition
4. **The oxidation number of oxygen in peroxides:** (LHR. GI, MLN. GII)
   (A) Zero (B) −1 (C) −2 (D) +2
5. **The oxidation number of oxygen in OF_2 is:** (LHR. GII, FBD. GI)
   (A) −1 (B) −2 (C) +1 (D) +2
6. **The oxidation state of sulphur in H_2SO_4 is:** (GRW. GI, BWP. GI)
   (A) +3 (B) +5 (C) +6 (D) −6
7. **Oxidation number of hydrogen in HCl is:** (GRW. GII)
   (A) +1 (B) −1 (C) 0 (D) −2
8. **The oxidation number of hydrogen in metal hydrides is:** (FBD. GI, SWL. GII)
   (A) 0 (B) −1 (C) +1 (D) −2
9. **The oxidation number of Chromium in $K_2Cr_2O_7$ is:** (MLN. GI, BWP. GI)
   (A) +2 (B) +6 (C) +7 (D) +14
10. **The oxidation number of chlorine in $KC\ell O_3$ is:** (SGD. GI, DGK. GI, DGK. GII)
    (A) +6 (B) +5 (C) +1 (D) −2
11. **Oxidation number of free element is:** (BWP. GII, SGD. GII)
    (A) 0 (B) +1 (C) +2 (D) +3

12. A specie that reduces a substance by donating electrons to it is called: (GRW. GII)

(A) oxidizing agent (B) reducing agent
(C) colouring agent (D) dehydrating agent

13. Non-metals act as oxidizing agents because: (GRW. GII)

(A) they are more electropositive (B) they are more electronegative
(C) they are neither electropositive nor electronegative
(D) they have low value of ionization energy

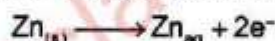
Answers

1.	Oxidation	2.	Electrochemistry			3.	Reduction	4.	−1
5.	+2	6.	+6	7.	+1	8.	−1	9.	+6
10.	+5	11.	0	12.	reducing agent	13.	they are more electronegative		

☆ Give short answer to the following questions.

1. Define oxidation in terms of electron and give an example.
(LHR. GII, RWP. GII, BWP. GI, SGD. GII, RWP. GII)

Ans. Loss of electron from an atom or ion, is called oxidation.

$$Zn_{(s)} \longrightarrow Zn_{aq} + 2e^-$$

2. Define oxidation and reduction. (SGD. GI, MLN. GII, GRW. GII, FBD. GI, BWP. GII)

Ans. During a chemical reaction, addition of oxygen and removal of hydrogen is called oxidation, while during chemical reaction, removal of oxygen and gain of hydrogen is called reduction.

3. Define oxidation number with an example. (SWL. GI, DGK. GII)

Ans. The number which represent the charge on an atom or ion is called oxidation number. For example in HCl there is +1 oxidation number of H and -1 for Cl.

4. Differentiate between spontaneous and non spontaneous reactions.
(SGD. GI, SGD. GII, GRW. GII, BWP. GI)

Ans. **Spontaneous reaction:** The reactions which can take place by their own without any external agent, are called spontaneous reaction.

Non-spontaneous reaction: The reaction which can not takes place by their own are called non-spontaneous reaction.

5. Define redox reaction. (RWP. GII, LHR. GI, LHR. GII)

Ans. Those chemical reaction in which oxidation as well reduction reaction has been take place are collectively called redox reaction.

6. What type of reaction takes place at anode in electrolytic cell? (BWP. GII)

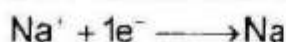
Ans. In electrolytic cell, there is oxidation reaction on anode.

7. Define oxidation and reduction in terms of loss and gain of electrons. (MLN. GI)

Ans. **Oxidation:** The removal of electron from an atom or ion is called oxidation.

e.g $Na \longrightarrow Na^{+} + 1e^{-}$

Reduction: Gain of electrons in an atom or ion is known as reduction.

$Na^{+} + 1e^{-} \longrightarrow Na$

8. What is spontaneous reaction? In which cell this reaction takes place? (MLN. GI)

Ans. The chemical reaction which takes place in the presence of external agent, is known as spontaneous reaction.

These chemical reaction occured in galvanic and electrolytic cell.

9. Define electrochemistry. Explain it with a reaction. (SWL. GI, DGK. GII)

Ans. A branch of chemistry that deals with relationship between electricity and chemical reaction. In this branch, redox reactions are studied.

$2ZnO + C \longrightarrow 2Zn + CO_2$

10. Define oxidation reaction. (SWL. GII)

Ans. During a chemical reaction, gain of oxygen and loss of hydrogen is known as oxidation.

11. Calculate oxidation number of nitrogen in HNO_3. (LHR. GII, LHR. GII, 2015)

Ans. The sum of oxidation numbers of a compound is zero. According to formula in HNO_3.

[oxidation number of H] + [oxidation number of N] +3 [oxidation number of O]=0

By putting values

[+1] + [oxidation number of N] +3 [–2] = 0

+1 + oxidation number of N + [–6] = 0

oxidtion number of N = 6 – 1 = 5

12. What is the difference between valency and oxidation state?

Ans. (GRW. GI, MLN. GI, SWL. GII, GRW. GII, SWL. GI)

Valency	Oxidation State
The combining capacity of an element with other element is called valency.	The number which represent the apparent charge on an atom or ion, known as oxidation state.

13. Calculate oxidation number of chlorine in $KClO_3$. (GRW. GII, DGK. GI)

Ans. [oxidation number of K] + [oxidation number of Cl]+3[oxidation number of O] = 0

[+1] + [oxidation number of Cl] +3[–2] = 0

+ 1 + [oxication number of Cl] +[–6] = 0

oxidation number of Cl = 6 – 1 = +5

14. Calculate oxidation number of sulphur in H_2SO_4 as (H = +1, O = - 2)

(FBD. GI, SGD. GI, BWP. GI, SWL. GI, RWP. GII)

Ans. 2[oxidation number of H]+[oxidation number of s]+4[oxidation number of O] = 0

2[+1] + [oxidation number of S] + 4 [–2] = 0

2 + [oxidation number of s] + [–8] = 0

oxidation number of S = 8 – 2

= 6

15. Calculate the oxidation number of Mn of $KMnO_4$. (BWP. GII)

Ans. (oxidation number of K)+(oxidation number of Mn)+4(oxidation number of O) = 0

1 + Mn + 4 (–2) = 0

1 + Mn + (–8) = 0

Mn = + 8 – 1

M = 7

16. Define reducing agent, also give one example. (RWP. GI, SWL. GII)

Ans. Those species which reduce other substance, and themselves get oxidize is known as reducing agent. e.g. All metals are good reducing agents.

7.4 Oxidation - Reduction Reactions

7.5 Electrochemical Cells

7.6 Electrochemical Industries

7.7 Corrosion and Its Prevention

☆ Tick the correct answer.

1. Formation of water from Hydrogen and Oxygen is a ________ reaction. (MLN. GI, DGK. GII, RWP. GI, MLN. GI)

(A) Redox (B) Acid-Base (C) Neutralization (D) Decomposition

2. In the redox reaction, between Zn and HCℓ the oxidizing agent is: (SWL. GI, BWP. GII, BWP. GI, DGK. GI, GRW. GI)

(A) Zn (B) H^+ (C) $C\ell$ (D) H_2

3. Which of the followings is oxidation reaction"? (DGK. GI)

(A) $K^+ + 1e \longrightarrow K$ (B) $Cl_2 + 2e^- \longrightarrow 2Cl$

(C) $Fe^{+2} \longrightarrow Fe^{+3} + 1e$ (D) Both A and B

4. A process of oxidation is: (SWL. GI)

(A) removal of oxygen (B) gain of electrons

(C) loss of electrons (D) addition of hydrogen

5. Which of the following electrolyte produces less ions in water: (LHR. GI, MLN. GII)

(A) H_2SO_4 (B) NaOH (C) $Ca(OH)_2$ (D) NaCℓ

6. In which cell spontaneous chemical reaction takes place? (GRW. GI)

(A) Electrolytic cell (B) Galvanic cell (C) Nelson's cell (D) Downs cell

7. **Types of electrochemical cell are:** (GRW. G II)
(A) 2 (B) 3 (C) 4 (D) 5

8. **Is a non electrolyte:** (FBD. GII)
(A) $HC\ell$ (B) $NaOH$ (C) C_6H_6 (D) H_2SO_4

9. **Which one of the following is not an electrolyte?** (SWL. GI, FBD. GI)
(A) sugar solution (B) sulphuric acid solution
(C) lime solution (D) sodium chloride solution

10. **Which one of the following is weak electrolyte:** (LHR. GII, SWL. GI)
(A) $NaC\ell$ (B) $NaOH$ (C) H_2SO_4 (D) CH_3COOH

11. **Pure water is an example of:** (LHR. GII)
(A) Weak electrolyte (B) Strong electrolyte (C) Strong acid (D) Strong base

12. **Which one of the following is strong electrolyte?** (SWL. GII)
(A) common salt solution (B) sugar solution
(C) pure water (D) benzene

13. **Anode of Downs cell is made of:** (LHR. GII)
(A) Steel (B) Copper (C) Zinc (D) Graphite

14. **What is obtained from fused NaCl?** (SGD. GI)
(A) NaOH (B) Sodium metal (C) Both A and B (D) None

15. **Which one of following method is used for production of sodium metal:** (SGD. GII)
(A) nelson's cell (B) downs cell (C) galvanic cell (D) electroplating

16. **Percentage of Cu is present in sterling silver:** (RWP. GII)
(A) 6.5 (B) 7.8 (C) 7.5 (D) 7.4

17. **The process of coating thin layer of zinc on iron is called:** (FBD. GI)
(A) Oxidizing (B) Reducing (C) Galvanizing (D) Alloying

18. **The formula of rust is:** (SWL. GII, SGD. GII, BWP. GI, LHR. GI)
(A) $Fe_2O_3.nH_2O$ (B) Fe_2O_3 (C) $Fe(OH)_3.nH_2O$ (D) $Fe(OH)_3$

19. **The most common example of corrosion is:** (RWP. GI, RWP. GII, FBD. GII)
(A) Chemical decay (B) Rusting of Iron
(C) Rusting of Aluminium (D) Rusting of tin

Answers

1.	Redox	2.	H^+	3.	$Fe^{+2} \longrightarrow Fe^{+3} + 1e^-$	4.	loss of electrons
5.	$Ca(OH)_2$	6.	Galvanic cell	7.	2	8.	C_6H_6
9.	sugar solution	10.	CH_3COOH	11.	Weak electrolyte	12.	common salt solution
13.	Graphite	14.	Sodium metal	15.	downs cell	16.	7.5
17.	Galvanizing	18.	$Fe_2O_3.nH_2O$	19.	Rusting of Iron		

☆ **Give short answer to the following questions.**

1. **Differentiate between oxidizing and reducing agents.** (LHR. GI, BWP. GI)

Ans. **Oxidizing agent:** The type of specie, that oxidize the substance by taking electron from it, and themselves get reduced by gaining the electrons e.g non-metals are oxidizing agents.

Reducing Agent: The type of specie that reduce other substances by giving them electrons, and themselves get oxidize. Almost all metals are good reducing agents, because they have tendency to lose electrons.

2. **Which substance is oxidized and reduced in following reaction. Identify these $2Na + Cl_2 \longrightarrow 2NaCl$.** (DGK. GI)

Ans. $2NaCl \longrightarrow 2Na + Cl_2$

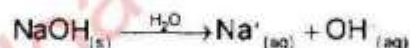
In this reaction sodium get oxidize while chlorine become reduce.

3. **Define electrolytic cell.** (LHR. GII, FBD. GI, GRW. GI, SWL. GI)

Ans. The type of electrochemical cell in which a non-spontaneous chemical reaction take place when electric current is passed through solution is called electrolytic cell.

4. **Differentiate between strong and weak electrolytes.** (FBD. GII, SGD. GI)

Ans. Strong electrolytes ionize completely in an aqueous solution. NaCl, NaOH and H_2SO_4 are example.

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

Weak electrolytes do not completely ionized in aqueous solution CH_3COOH and $Ca(OH)_2$ are examples.

$$CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

5. **Explain Non-electrolytes with an example.** (MLN. GII, BWP. GII, LHR. GI)

Ans. Those species that do not ionize in an aqueous solution, and current is not passing through their solution is called non-electrolyte. e.g. sugar solution.

6. **What is the difference between electrolytic cell and galvanic cell?** (SGD. GI, MLN. GII)

Ans. **Electrolytic cell:**

i. It consist of complete cell, which is connected with battery.
ii. Anode has positive while cathode has negative charge.
iii. Electrical energy is converted into chemical energy.
iv. Current is used for non-spontaneous chemical reaction.

Galvanic cell:

i. It consist of two half cells.
ii. Anode has negative while cathode has positive charge.
iii. Chemical energy is converted into electrical energy.

iv. Redox reaction takes place by its own.

7. What is salt bridge? What is its basic function? (SGD. GII, GRW. GI, DGK. GII)

Ans. The half cells of galvanic are connected by salt bridge. It provide pathway for migration of ions. And keep the solution neutral.

8. What is meant by electrolysis? (RWP. GII)

Ans. When an electric current is passing through a aqueous solution or melted form of a compound, it decomposes into its basic component, this process is known as electrolysis.

9. What is electrochemical cell? Write the name of its types. (DGK. GI, FBD. GII)

Ans. A type of cell in which two electrodes are immersed in electrolyte solution and both are connected with battery.
There are two types of electrochemical cell:
Electrolytic cell, Galvanic cell.

10. What happens at the Cathode in a Galvanic Cell? (BWP. GI)

Ans. In galvanic cell, reduction has been done.

11. Name the By - Products produced in Nelson's Cell. (BWP. GI, SGD. GII)

Ans. In Nelson cell, two by products hydrogen and chlorine are formed.

12. Where do the electrons flow from Zn electrode in Daniel's cell? (BWP. GII)

Ans. In Daniel cell, electrons accumulate on zinc electrode, which moves toward cathode through external circuit.

13. Write the names of Electro chemical cells. (LHR. GI, SGD. GII)

Ans. There are two types of electrochemical cell.
(i) Electrolytic cell
(ii) Galvanic cell.

14. What is the difference between Anode and Cathode? (LHR. GI)

Ans. That electrode which is connected to positive terminal known as anode. That electrode which is connected to negative terminal is known as cathode.

15. How the half cells of galvanic cell are connected? What is the function of salt bridge? (LHR. GII)

Ans. The half cells are connected through salt bridge. The bride provide the pathway for migration of ion in solution.

16. Are anodes of Downs cell and Nelson's cell made of the same element? If yes, write its name. (GRW. GI)

Ans. Yes, the anode of down cell and nelson cell are made up of same element, graphite.

17. Define Galvanic cell and give one example. (SWL. GI, DGK. GI)

Ans. That type of electrochemical cell, in which spontaneous chemical reaction take

place and current is produced.

e.g Daniel cell.

18. Define electrolytes, also give one example. (SWL. GII)

Ans. Those substances which allow to pass electricity through their solution or molten form is known as electrolyte.

e.g Solution of acid, bases and salts are good electrolytes.

19. Identify the electrolyte and non electrolyte from the following:

(i) Sugar (ii) Glucose (iii) Benzene (iv) Sodium chloride (RWP. GI)

Ans. Electrolyte = Glucose, Sodium chloride

Non-electrolyte = Sugar, Benzene.

20. Write down names of any two weak electrolyte. (DGK. GI)

Ans. 1. CH_3COOH

2. $Ca(OH)_2$

21. Why galvanizing is done? (GRW. GI, SWL. GII, GRW. GI, MLN. GI, SGD. GI)

Ans. To prevent from rust, iron is being galvanized.

22. What is meant by alloying? (FBD. GI, GRW. GII)

Ans. The homogenous mixture of metal with other metal or non-metal is called alloying. Alloy of iron with other metal prove to be successful technique against rust pollution. The good example of its stean less steal. Which is mixture of iron, chromium and nickel.

23. Define metallic coating. In which industry it is used much? (SGD. GI, MLN. GI)

Ans. The best method to prevent metals, from being rusting is metallic coating. Tin, Zinc and chromium are used in coating. This technique is commonly used in food industry. Where food is preserved in boxes, these iron boxes are coated with tin or chromium.

24. In electroplating of chromium, which salt is used as electrolyte? (LHR. GI)

Ans. In this process the salt of chromium sulphate is used as an electrolyte.

25. Why O_2 is necessary for rusting? (LHR. GI, RWP. GI, SGD. GI, FBD. GII)

Ans. Because oxygen act as oxidizing agent. It accepts electron from iron and change it into ferrous ion (Fe^{2+}) and then ferric ion (Fe^{3+}). Oxygen combine with ferrous ion in the presence of water and form rust ($Fe_2O_3.nH_2O$).

26. Why an iron grill is painted frequently? (GRW. GI)

Ans. In order to prevent from rusting, the iron grill painted frequently. Because due to moisture in air, it is at risk of being rusted.

27. Define electroplating. (FBD. GII, LHR. GI, FBD. GII)

Ans. The layering of one metal to another through electrolysis is called electroplating.

28. Define corrosion with an example. (SWL. GI, DGK. GII)

Ans. It is slow and continuous process of eating away of a metal by the surrounding medium. It is a redox reaction, that takes place by the action of air and moisture with metals.

29. What do you mean by rusting of iron? (SWL. GII, FBD. GI)

Ans. A slow and continuous process of eating away of metal by its surrounding, called rusting of iron.

30. How electroplating of Tin on steel is carried out? (DGK. GI, SGD. GI)

Ans. Tin is usually electroplated on steel by placing the steel into a container containing a solution of tin salt. The steel is connected to a electrical circuit, acting as cathode. While other electrode made of tin metal act as anode. When an electrical current passes through circuit, tin metal ion present in solution deposit on steel.

31. What is the shape of cathode in Nelson's cell? Why is it perforated? (BWP. GII)

Ans. In Nelson cell, cathode contain a perforated tank. It is perforated because sodium hydroxide solution has to percolates into catch basin.

32. What is the difference between corrosion and rusting? (BWP. GII)

Ans. Continuously eating away of metal by the result of oxidation is known as corrosion, while corrosion of iron is known as rusting.

33. How impure copper is refined to pure copper? (FBD. GI)

Ans. Copper is refined through electrolytic process impure copper act as anode and pure copper acts as cathode. The solution of copper sulphate act as electrolyte.

The oxidation is carried out on anode. Impure copper gives electrons to anode, and dissolves in solution as copper ion.

The process of reduction is carried on cathode. Copper ion attract toward cathode and attain electron from cathode.

$Cu^{+2} + 2e^{-} \longrightarrow Cu$

34. How electroplating of Zinc is carried out? (SWL. GI)

Ans. The target metal is cleaned in alkaline detergent type solution, and it is treated with acid, in order to remove any rust or surface scaller. Next the zinc is deposited on the metal by immersing it in a chemical bath, containing electrolyte zinc sulphate. A current is applied, which results in zinc being deposited on targeted metal.

35. Name the products obtained from downs' cell. (DGK. GI)

Ans. Sodium metal is the compound, which is gain in down cell.

CHEMICAL REACTIVITY

Major Concepts:

8.1 Metal

8.2 Non-Metals

Time allocation
Teaching periods 07
Assessment periods 02
Weightage 10%

Students Learning Outcomes:
Students will be able to:

- Show how cation and anion are related to the terms metals and non-metals.
- Explain Alkali metals are not found in the free state in nature.
- Explain the differences in ionization energies of alkali and alkaline earth metals.
- Describe position of sodium metal in the periodic table its simple properties and uses.
- Position of calcium and magnesium in the periodic table, their simple properties and uses.
- Differentiate between soft and hard metals (iron and sodium)
- Describe the inertness of noble metals.
- Identify commercial value of silver, gold and platinum.
- Compile some important reactions of halogens.
- Name some elements that exist in nature in uncombined form.

8.1 METALS

Q.1(a) Define chemical reactivity.
(b) What are metals? Write down their classification.
(c) Write down some physical and chemical properties of metals.

Ans. (a) Chemical reactivity: The tendency of an element to react with another element is called chemical reactivity. It is measured by the relative tendency of an element to lose or gain electrons in chemical reactions.

(b) Metals: Metals are the elements (except hydrogen) which are electropositive and form cations by losing electrons.

Types of metals: Metals are classified into three types.

1. **Very reactive:** Potassium, sodium, calcium, magnesium and aluminium.
2. **Moderately reactive:** Zinc, iron, tin and lead
3. **Least reactive or noble:** Copper, mecury, silver and gold.

(c) Physical properties of metals:

Some important physical properties of metals are given as follows:

Some common metals and non-metals.

(i) Almost all metals are solids (except mercury)

(ii) They have high melting and boiling points.

(iii) They possess metallic luster and can be polished.

(iv) They are malleable (can be hammered into sheets), ductile (can be drown into wires) and give off a tone when hit.

(v) They are good conductor of heat and electricity.

(vi) They have high density.

(vii) They are hard (except sodium and potassium)

Chemical properties of metals:

Some important chemical properties of metal are following.

(i) They easily lose electrons and form positive ions.

(ii) They readily react with oxygen to form basic oxides.

(iii) They usually form ionic compounds with non-metals.

(iv) They have metallic bonding.

Do you know?

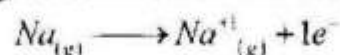
☆ The most abundant metal is aluminium
☆ The most precious metal is platinum
☆ The most useable metal is iron
☆ The most reactive metal is cesium
☆ The most valuable metal is uranium
☆ The lightest metal is lithium (d=0.53 g cm^{-3})
☆ The heaviest metal is osmium (d = 22.5 g cm^{-3})
☆ The least conductor of heat is lead.
☆ The best conductor metals are silver and gold
☆ The most ductile and malleable metals are gold and silver

Q.2.(a) What is meant by electropositive character? Discuss the trends of electropositivity in the periodic table.

(b) Discuss the relationship between electropositivity and ionization energy.

Ans.(a) Electropositive character/electropositivity: The tendency of an element (metal) to lose one or more electrons to form a positive ion is called electropositive character or electropositivity.

Explanation: Metals have high electropositive character because they readily lose electrons to form positive ions [due to low ionization energies]. e.g.

$$Na_{(g)} \longrightarrow Na^{+1}_{(g)} + 1e^-$$

Trends along periodic table:

Trends along group: Electropositive character [electropositivity] increases down the group due to increase in atomic size.

Example: Lithium metal is less electropositive than sodium which is less electropositive than potassium.

Trends along period: Electropositive character decreases from left right in a period due to decrease in atomic size and increase in nuclear charge.

(b) Electropositivity and ionization energy:

Electropositivity (electropositive character) character depends upon ionization energy, which in turn depends on size and nuclear charge of the atom.

Small size atoms with high nuclear charge have high ionization energy, hence atoms having high ionization energy are less electropositive or less metallic.

Alkali metals have the largest size and the lowest ionization energy in their respective periods. Therefore they have the highest metallic character.

Metal	Atomic Number	Electronic Configuration	IE	Metal	Atomic Number	Electronic Configuration	IE_1	IE_2
Li	3	$[He]2s^1$	520	Be	4	$[He]2s^2$	899	1787
Na	11	$[Ne]3s^1$	496	Mg	12	$[Ne]3s^2$	738	1450
K	19	$[Ar]4s^1$	419	Ca	20	$[Ar]4s^2$	590	1145
Rb	37	$[Kr]5s^1$	403	Sr	38	$[Kr]5s^2$	549	1064
Cs	55	$[Xe]6s^1$	377	Ba	56	$[Xe]6s^2$	503	965

Test yourself 8.1:

i. **What type of elements are metals?**

Ans. Metals are the elements (except hydrogen) which are electropositive and form cations by losing electrons.

ii. Name a metal which exists in liquid form?
Ans. Mercury is a metal which exists in liquid form.
iii. What is the nature of metal oxide?
Ans. Mostly the metal oxides are basic in nature.
iv. Which group of metals is highly reactive?
Ans. Group I of metals is highly reactive.
v. Why sodium metal is more reactive than magnesium metal?
Ans. Sodium is more reactive than magnesium because the ionization energy of sodium is less than that of magnesium.
vi. Name a metal which can be cut with knife?
Ans. Sodium is a metal which can be cut with knife.
vii. Name the best ductile and malleable metal?
Ans. Gold and silver
viii. Name the metal which is the poorest conductor of heat?
Ans. Lead is the poorest conductor of heat.
ix. What do you mean by malleable and ductile?
Ans. **Malleable:** Can be converted into sheet.
Ductile: Can be converted into wires.
x. Why alkali metals are more reactive than alkaline earth metals?
Ans. Alkali metals are more reactive than alkaline earth metals because the ionization energies of alkali metals are less than alkaline earth metals.
xi. What do you mean by metallic character?
Ans. The ability of metals to lose electrons is called metallic character.
xii. Why metallic character decreases along a period and increases in a group?
Ans. Metallic character increases down the group due to increase in atomic size and decrease in nuclear force.
The metallic character decreases from left to right in a period due to increase in nuclear force and decrease in atomic size.

Q.3 Compare the physical properties of alkali and alkaline earth metals by giving the comparison of sodium, magnesium and calcium.

Ans. Comparison of Physical Properties of Alkali and Alkaline Earth Metals

Property	Sodium	Magnesium	Calcium
Appearance	Silvery white having a metallic luster, very soft and can be cut with knife	Silvery white and hard	Silvery grey and fairly harder
Atomic size, ionic size (pm)	186,102	160,65	197,99
Relative density	0.98g cm^{-3} Floats on water	1.74gcm^{-3}	1.55gcm^{-3}
Malleability	very malleable and ductile	malleable and ductile	malleable and ductile
Conductivity	Good conductor of heat and elecricity	Good conductor of heat and electricity	Good conductor of heat and electricity
M.P	97°C	650°C	839°C

M.P	883°C	1090°C	1484°C
Ionization energy	496kJ/mol	738,1450 kJ/mol	590,1145 kJ/mol
Flame in air	Golden yellow	Brilliant white	Brick red

Q.4. Compare the chemical properties and chemical reactivities of alkali and alkaline earth metals.

Ans. A comparison of chemical properties and reactivities of alkali metals and alkaline earth metals is given below.

Alkali Metals	Alkaline Earth Metals
1. Occurrence: They are very reactive and always occur in combined form.	They are fairly reactive and also occur in combined form.
2. Electropositivity: These are highly electropositive. They have ionization energy values ranging from 520kJmol^{-1} for Li to 376kJmol^{-1} for Cs.	They are less electropositive. They have ionization energy values ranging from 1757kJmol^{-1} for Be to 965 kJmol^{-1} for Ba.
3. Reaction with water: They react with water vigorously at room temperature to give strong alkaline solution and hydrogen gas $2Na+2H_2O \longrightarrow 2NaOH+H_2$	They react with water less vigorously and on heating they produce weak bases $Mg+H_2O \longrightarrow MgO + H_2$ $MgO+H_2O \longrightarrow Mg(OH)_2$
4. Reaction with O_2: They immediately tarnish in air giving their oxides which form strong alkalies in water. $4Na+O_2 \longrightarrow 2Na_2O$ $Na_2O+H_2O \longrightarrow 2NaOH$	They are less reactive towards oxygen and oxides are formed on heating $2Mg+O_2 \longrightarrow 2MgO$
5. Reaction with Hydrogen: They form ionic hydrides with H_2 at high temperature $2Na+H_2 \longrightarrow 2NaH$	They give hydride under strong conditions of temperature and pressure. $Ca+H_2 \longrightarrow CaH_2$
6. Reaction with Halogens: They react violently with halogens at room temperature to give halides $2Na+Cl_2 \longrightarrow 2NaCl$	They react slowly with halogen to give their halides $Ca+Cl_2 \longrightarrow CaCl_2$
7. Reaction with Nitrogen: They do not form nitrides directly.	They form stable nitrides when heated with nitrogen. $3Mg+N_2 \longrightarrow Mg_3N_2$
8. Reaction with carbon: They do not react with carbon directly.	They give stable carbide on heating with carbon. $Ca+2C \longrightarrow CaC_2$

Q.5. Write down the use of sodium, magnesium and calcium.

Ans. Uses of Sodium:

1. Sodium -potassium alloy is used as a coolant in nuclear reactors.
2. It is used to produce yellow light in sodium vapour lamps.
3. It is used as a reducing agent in the extraction of metal like Ti.

Uses of magnesium:

1. Magnesium is used in flash lights and in fireworks.
2. It is used in the manufacture of light alloys.
3. Magnesium ribbon is used in Thermite process to ignite aluminium powder.
4. Magnesium is used as anode for prevention of corrosion.

Uses of calcium:

1. It is used to remove sulphur from petroleum products.
2. It is used as reducing agent to produce Cr, U and Zr.

Q.6. Discuss the inert character of silver, gold and platinum.

Ans. Inert character of Silver: Silver is white lustrous metal. It is an excellent conductor of heat and electricity. It is also highly ductile and malleable metal. Its polished surfaces are good reflectors of light. Formation of thin layer of oxide or sulphide on its surface makes it relatively un-reactive. Under normal conditions of atmosphere, air does not affect silver. It tarnishes in presence of sulphur containing compounds like H_2S. Being very soft metal, it is rarely used as such. Alloys of silver with copper are widely used in making coins silver-ware and ornaments. Compounds of silver are widely used in photographic films and dental preparations. Silver also has important applications in mirror industry.

The Transition Elements in the Periodic Table

Inert Character of Gold: Gold is a yellow soft metal. It is most malleable and ductile of all the metals. One gram of gold can be drawn into a wire of one and a half kilometer long. Gold is very non-reactive or inert metal. It is not affected by atmosphere. It is even not affected by any single mineral acid or base. Because of its inertness in atmosphere, it

is an ornamental metal as well as used in making coins. Gold is too soft to be used as such. It is always alloyed with copper, silver or some other metal.

Do you know?

Purity of gold is shown by carats that indicates the number of parts by weight of gold that is present in 24 parts of alloy. Twenty four carat gold is pure. 22 carats gold means that 22 parts pure gold is alloyed with 2 parts of either silver or copper for making ornaments and jewelry. White gold is its alloy with palladium, nickel or zinc.

Inertness of Platinum: Platinum is used to make jewelry items because of its unique characteristics like colour, beauty, strength, flexibility and resistance to tarnish. It provides a secure setting for diamonds and other gemstones, enhancing their brilliance. Platinum alloy with palladium and rhodium is used as catalyst in auto-mobiles as catalytic convertor. They convert most of the toxic gases being emitted by vehicles into less harmful carbon dioxide, nitrogen and water vapour. Platinum is used in the production of hard disk drive coatings and fibre optic cables. Platinum is used in the manufacturing of fibre glass reinforced plastic and glass for liquid crystal displays (LCD).

Q7. (a) What are noble metals? Give examples. Indicate their position in the periodic table.

(b) Define alkali and alkaline earth metals with examples.

Ans.(a) Noble metals: The metals which are relatively inactive, do not lose electrons easily, highly resistant to oxidation and corrosion are called noble metals.

Examples: Gold, platinum and silver.

Position in Periodic Table: They are found among the heavier transition elements which are present in the center of periodic table.

(b) Alkali metals: The metals of group IA of the periodic table are called alkali metals e.g. Lithium, sodium, potassium etc.

Alkali metals are extremely reactive because they have only one electron in their last shell, and their valence shell electronic configuration is ns^1. They readily form salts with non-metals.

Alkaline earth metals: The elements which are present in group IIA of the periodic table are called alkaline earth metals, e.g. calcium, magnesium etc. They have two electrons in their last shells. They are also reactive but less than alkali metals.

Test yourself 8.2:

i. **Give the applications of silver?**

Ans. Silver is used in making ornaments. Alloys of silver are used in making coins, silver-wares. The compounds of silver are widely used in photographic films and dental preparation. It is also used in mirror industry.

ii. **Why silver is not used in pure form?**

Ans. Silver metal is not used in pure form due to its softness.

iii. **What do you mean by 24 carat gold?**

Ans. Purity of gold is shown by carats that indicates the number of parts by weight of gold that is present in 24 parts of alloy. Twenty four carat gold is pure.

iv. Why gold is used to make jewelry?

Ans. The gold is used to make jewelry due to its inertness in atmosphere.

v. Why platinum is used for making jewelry?

Ans. Platinum is used to make jewelry items because of its unique characteristics like colour beauty, strength, flexibility and resistance to tarnish.

vi. What is difference between steel and stainless steel?

Ans. **Stainless steel:** It contain iron, nickel and chromium.

Steel: Simple steel contains inon and carbon.

vii. How platinum is used as a catalyst in automobiles and what are advantages of this use?

Ans. Platinum alloy with palladium and rhodium is used as catalyst in automobiles as catalytic converter.

Advantage: It converts most of the gases being emitted by vehicles into less harmful Carbon dioxide, nitrogen and water vapour.

8.2 NON-METALS

Q8. Write down some physical and chemical properties of non-metals.

Ans. Definition of non-metals: Non-metals from negative ions (axons) by gaining electrons. In this way non-metals are electronegative in nature and form acidic oxides.

Important physical properties of non-metals: Some important physical proper ties of non-metals are given below.

1. Solids non-metals are brittle (break easily).
2. Non-metals are non-conductor of heat and electricity (except graphite).
3. They are not shiny, they are dull except iodine (it is lustrous like metals).
4. They are generally soft (except diamond).
5. They have low melting and boiling points (except silicon, graphite and diamond).
6. They have low densities.

The Non-Metals in Periodic Table

Important chemical properties of non-metals:

1. Their valence shells are deficient of electrons, therefore they readily accept electrons to complete their valence shells and become stable.
2. They form ionic compounds with metals and covalent compounds by reacting with other non-metals e.g. CO_2, NO_2, etc.
3. Non-metals usually do not react with water.
4. They do not react with dilute acids because non-metals are itself electron acceptors.

Q9. Discuss the significance of Non-metals.

Ans. Significance of non-metals: Although non-metals are fewer than metals, yet they are highly significant. They are equally important for human beings, animals and plants. In fact, life would not have been possible without the presence of non-metals on earth.

1. Major components of earth's crust, oceans and atmosphere are non-metals: oxygen has the highest percentage in earth's crust (47%) and oceans (86%) and it is second (21%) to nitrogen in atmosphere. It indicates the importance of oxygen in nature. To maintain the balance for the amount of non-metals in nature different cycles like water cycle, nitrogen cycle etc have been established naturally.
2. Non-metals are essential part of the body structure of all living things. Human body is made up of about 28 elements. But about 96% of the mass of the human body is made up of just 4 elements i.e. oxygen 65% carbon 18% hydrogen 10% and nitrogen 3% similarly plant bodies are made up of cellulose, which is composed of carbon, hydrogen and oxygen.
3. Life owes to non-metals as without O_2 and CO_2 (essential gases for respiration of animals and plants respectively), Life would not have been possible. In fact, these gases are essential for the existence of life.
4. All eatables like carbohydrates, proteins, fats, vitamins, water, milk etc which are necessary for the growth and development of body are made up of non-metals; carbon, hydrogen and oxygen. Its shows non-metals play a vital role for the maintenance of life.
5. The essential compound for the survival of life of both animals and plants is water, which is made up of non-metals. Water is not only major part by mass of animals and plants bodies, but it is also essential to maintain the life. We can survive without water for days but not for a long period; its shortage may cause death.
6. Another important non-metal is nitrogen, which is 78% in atmosphere, is necessary for the safety of life on earth. It controls the fire and combustion processes, other wise all the things around us could burn with a single flame.
7. Non-metals are playing essential role for the communication in life. All fossil fuels which are major source of energy; coal petroleum and gas are made up of carbon and hydrogen. Even the essential component of fossil fuels, oxygen is also a non-metal.
8. Non-metals protect us in a way, the clothes we wear are made of cellulose (natural fiber) or polymer (synthetic fiber).
9. In addition to all of these, other items used in daily life such as wooden or plastic furniture, plastic sheets and bags, plastic pipes and utensils are made of non-metals as major constituents.

Q10. What are non-metals? Briefly discuss the non-metallic character of elements.

Ans. Non-metals: The elements which form negative ions (anions) by gaining electrons

are called non- metals.

These are present in the upper right hand portion of the periodic table.

Examples: Some common examples of non-metals are following.

(i) Nitrogen (ii) Oxygen (iii) Phosphorous (iv) Sulphur (v) Bromine

Non-metallic character OR Electronegative character: The tendency of the elements to accept one or more electrons to form negative ions (anions) is called non-metallic character or electronegative character.

Explanation: The non-metallic character depends upon the electron affinity and electronegativity of the atom. Small size atoms having high nuclear charge are electronegative in nature.

They have high electron affinity and possess non-metallic character.

Trends along periodic table:

Along group:

Non-metallic character decreases down the group due to increase in atomic size.

Along period: Non-metallic character increases from left to right in a period due to decrease in atomic size.

Q.11. What are halogens? Indicate their position in the periodic table. Compare the properties of halogens.

Ans. Halogens:

The elements of group "17" of the periodic table are called halogens. they consist of

(i) Fluorine (ii) Chlorine (iii) Bromine (iv) Iodine (v) Astatine

Position in the periodic table: Halogens are present at right side of the periodic table.

Properties of Halogens:

1. Physical state of Halogens: Fluorine and chlorine are gases. Bromine is a liquid. Iodine is a solid. Astatine is a radioactive element.

2.Colour: All the halogens are coloured. The colour of halogens deepens down the group.

Fluorine is a pale yellow gas

Chlorine is a greenish yellow gas.

Bromine is a reddish brown liquid.

Iodine is a purple black solid

3. Solubility: Halogens are slightly soluble in water. their solubility decreases down the group.

Valence shell electronic configuration of halogens is ns^2 np^5.

All the Halogens are diatonic. They have seven electrons in their last shells and required only one electron to complete their last shells. Halogens form ionic bonds with metals and covalent bonds with non-metals.

Q.12. Describe the important reactions of halogens.

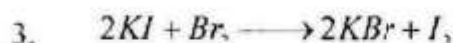
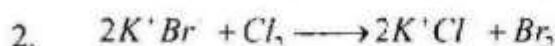
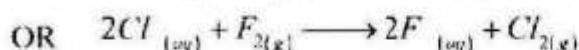
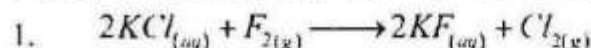
Ans. 1. Oxidizing properties: All the halogens are oxidizing agents. Fluorine is the

strong oxidizing element while iodine is the least. Fluorine will oxidize any of halide ion (X^-) in solution and changes to F^- ion. Chlorine will displace Br^- and I^- ion from their salts solutions and oxidize them to bromine and iodine. e.g.

1. $2KCl_{(aq)} + F_{2(g)} \longrightarrow 2KF_{(aq)} + Cl_{2(g)}$

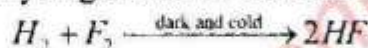
OR $2Cl^-_{(aq)} + F_{2(g)} \longrightarrow 2F^-_{(aq)} + Cl_{2(g)}$

2. $2K^+Br^- + Cl_2 \longrightarrow 2K^+Cl^- + Br_2$

3. $2KI + Br_2 \longrightarrow 2KBr + I_2$

2. Reaction with Hydrogen: All the halogens combine with hydrogen to give hydrogen halides. (HX).

The chemical affinity of halogens for hydrogen decreases from F_2 to I_2.

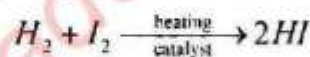
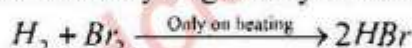
Fluorine combines with hydrogen even in the dark and cold state.

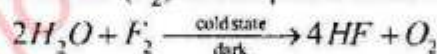
$$H_2 + F_2 \xrightarrow{\text{dark and cold}} 2HF$$

Chlorine combines with hydrogen in the presence of sunlight.

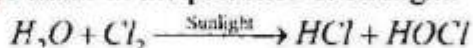
$$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$$

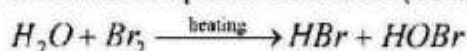
Bromine and iodine reacts with hydrogen only on heat.

$$H_2 + Br_2 \xrightarrow{\text{Only on heating}} 2HBr$$

$$H_2 + I_2 \xrightarrow[\text{catalyst}]{\text{heating}} 2HI$$

3. Reaction wih water: Fluorine (F_2) decomposes water in cold state in dark.

$$2H_2O + F_2 \xrightarrow[\text{dark}]{\text{cold state}} 4HF + O_2$$

Chlorine decomposes water in the presence of sunlight.

$$H_2O + Cl_2 \xrightarrow{\text{Sunlight}} HCl + HOCl$$

Bromine reacts with water under special condition (heating)

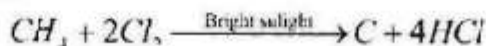
$$H_2O + Br_2 \xrightarrow{\text{heating}} HBr + HOBr$$

Iodine does not react with water.

$$H_2O + I_2 \longrightarrow \text{No reaction}$$

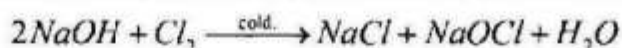
4. Reaction with one methane:

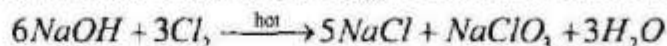
Fluorine (F_2) reacts violently with methane (CH_4) in dark.

Chorine (Cl_2) does not react with methane in dark. It reacts violently in the presence of bright sunlight.

$$CH_4 + 2Cl_2 \xrightarrow{\text{Bright sunlight}} C + 4HCl$$

In the presence of diffused sunlight, the reaction of chlorine with methane is slow and gives series of compounds. i.e. CH_3Cl, CH_2Cl_2, $CHCl_3$ and CCl_4.

5. Reaction with sodium hydroxide: Chlorine reacts with cold dilute sodium hydroxide to give sodium hypochlorite.

$$2NaOH + Cl_2 \xrightarrow{cold.} NaCl + NaOCl + H_2O$$

Chlorine reacts with hot conc. NaOH to give sodium chlorate and sodium chloride.

$$6NaOH + 3Cl_2 \xrightarrow{hot} 5NaCl + NaClO_3 + 3H_2O$$

Test yourself 8.3:

i. **Why valency of chlorine is 1?**

Ans. Chlorine has seven electrons in its valance shell so it can gain only one electron or share one electron to complete its octet hence its valancy is 1.

ii. **Which factor controls the non-metallic character of the elements?**

Ans. Electron affinity and electronegativity, controls the non-metallic character of elements.

iii. **Why fluorine is more non-metallic than chlorine?**

Ans. Fluorine is more reactive than chlorine because electronegativity of fluorine is greater than chlorine.

iv. **Iodine exists in solid state, can it be beaten with hammer to form sheets?**

Ans. No iodine can not be converted into sheets by hammering because it is a non-metals.

v. **Can liquids and gases be brittle?**

Ans. Yes, liquids and gases are brittle.

vi. **Why the oxygen is called non-metal?**

Ans. Oxygen is non-metal because it completes its valence shell by giving electrons. It exists in gaseous state at room temperature.

vii. **Name two non-metals which are both brittle and non-ductile.**

Ans. Carbon and iodine are both brittle and non-ductile.

viii. **Name the most abundant non-metal in the earth's crust.**

Ans. Oxygen is the most abundant non-metal in the earth's crust.

ix. **Give the non-metallic tread in halogens.**

Ans. Non-metallic character of Halogens decreases down the group due to increase in atomic size.

x. **Why do the non-metals accept electrons readily?**

Ans. Non- metals having high electronegativity so they can accept electrons readily inorder to complete their last shells.

xi. **Why non-metals do not react with dilute acids while metals do react?**

Ans. Non- metals do not react with acids (dil) because non-metals are itself electron acceptors. Metals react with dilute acids because they can lose electrons readily.

xii. **How can we distinguish a metal from a non-metal by simple physical methods?**

Ans.

Metal	Non-metal
Metals are good conductors of electricity. Metals have luster.	Non-metals are not good conductors of electricity except graphite. Non-metals have not luster except iodine

xiii. **How can we distinguish a substance is metal or non-metals with the help of an acid?**

Ans. The substance which reacts with dil acids is called metal and which does not react is called non-metal.

xiv. **Why is HF a weak acid?**

Ans. Acids are the substances which can give hydrogen ions in aqueous solutions. HF is a weak acid because it does ionize in aqueous appreciably due to the presence of strong hydrogen bonding in it.

Key Points

- Formation of cations of alkali and alkaline earth metals is due to their electropositive behavior.
- The chemical reactivity of alkali and alkaline earth metals, is quite different.
- Calcium and magnesium are less reactive than sodium.
- Halogens form very stable compounds with alkali metals.
- Mercury and gold exist in free elemental form in nature.

Exercise (Solved)

☆ **Multiple Choice Questions**

Put a (✓) on the correct answer.

1. **Metals can form ions carrying charges:**
   (a) Uni-positive (b) Di-positive (c) Tri-positive (d) All of them
2. **Which one of the following metal burns with a brick red flame?**
   (a) Sodium (b) Magnesium (c) Iron (d) Calcium
3. **Sodium is extremely reactive metal, but it does not react with:**
   (a) Hydrogen (b) Nitrogen (c) Sulphur (d) Phosphorus
4. **Which one of the following is the lightest metal?**
   (a) Calcium (b) Magnesium (c) Lithium (d) Sodium
5. **Pure alkali metals can be cut simply by knife but iron cannot because of alkali metals have:**
   (a) Strong metallic bonding (b) Weak metallic bonding
   (c) Non-metallic bonding (d) Moderate metallic bonding
6. **Which of the following is less malleable?**
   (a) Sodium (b) Iron (c) Gold (d) Silver
7. **Metals lose their electrons easily because:**
   (a) They are electronegative (b) They have electron affinity
   (c) They are electropositive (d) Good conductors of heat
8. **Which one of the following is brittle?**
   (a) Sodium (b) Aluminium (c) Selenium (d) Magnesium
9. **Which one of the following non-metal is lustrous?**
   (a) Sulphur (b) Phosphorus (c) Iodine (d) Carbon
10. **Non-metals are generally soft, but which one of the following is extremely hard?**
    (a) Graphite (b) Phosphorus (c) Iodine (d) Diamond
11. **Which one of the following will not react with dilute HC1?**
    (a) Sodium (b) Potassium (c) Calcium (d) Carbon

Answers:

1. All of them 2. Calcium 3. Nitrogen 4. Sodium 5. Weak metallic bonding
6. Sodium 7. They are electropositive 8. Sodium 9. Iodine
10. Diamond 11. Carbon

☆ Short Answer Questions.

1. Why reactivity of metals increases down the group?

Ans. The reactivity of metals increases down the group due to increase in atomic size down the group.

The ability of the metals to lose electrons increases down the group hence their reactivity increases down the group.

2. State the physical properties of metals.

Ans. For answer see Q. 1

3. Why nitrogen forms compounds with alkaline earth metals directly?

Ans. Nitrogen forms compounds with alkaline earth metals directly because they form stable nitrides when heated with nitrogen. e.g. $3Mg + N_2 \longrightarrow Mg_3N_2$

4. Why the second ionization energy of magnesium is higher than the first one?

Ans. The second ionization energy of magnesium is higher than the first because the removal of electron from Mg^+ ion is difficult as the nuclear charge attracts the remaining electrons strongly. The size also decreases which contributes the increase in 2nd ionization energy.

5. How oxygen reacts with group 2 metals?

Ans. Oxygen reacts with II A metals to give their oxides e.g. $2Mg + O_2 \xrightarrow{heat} 2MgO$

6. What is relationship between electropositivity and ionization energy?

Ans. For answer see Q. 2

7. Why electropositivity decreases from left to right in a period?

Ans. For answer see Q. 2

8. How electropositivity depends upon size and nuclear charge of an atom?

Ans. For answer see Q. 2

9. Why ionization energies of alkaline earth metals are higher than alkali metals?

Ans. The ionization energies of alkaline earth metals are higher than alkali metals because alkali metals are having only one electron in their last shell which is required to remove while alkaline earth metals have two electrons in their last shell which are required to remove, so 2nd ionization energy is also required which is always greater than first.

10. Why are silver and gold least reactive?

Ans. Silver and gold are least reactive because they do not lose electrons easily.

11. Can pure gold be used for making ornaments? If not why?

Ans. Pure Gold can not be used for making ornament because gold is too soft to be used

as such. Got is always alloyed with copper, silver or some other metal.

12. Why is copper used for making electrical wires?

Ans. Copper is used for making electrical wises because it is a good conductor of electricity and can be easily converted into wires.

13. What is the trend of variation in densities of alkali metals?

Ans. Densities of alkali metals decreases down the group.

14. Which metal is used for metal work?

Ans. All those metals are used in metal works which are less reactive and cheap. For this purpose copper metal is commonly used.

15. Why is magnesium harder than sodium?

Ans. Magnesium is harder than sodium because magnesium form stronger metallic sound than sodium.

16. Why is calcium more electropositive than magnesium?

Ans. Calcium is more electropositive than magnesium because calcium has greater size than magnesium and calcium has greater ability to lose the electron than magnesium.

17. Why is ionization energy of Na less than Mg?

Ans. The ionization energy of sodium is less than magnesium because sodium has only one electron in the last shell and less energy is required to remove it while magnesium has two electrons in last shell and larger energy is required to remove these electrons.

18. Why is the ionization energy of Na more than K?

Ans. The ionization energy of sodium is more than potassium because the atomic size of sodium is smaller than potassium hence greater energy is required to remove the electron from sodium than potassium.

Long Answer Questions

1. Compare and contrast the properties of alkali and alkaline earth metals.

Ans: For answer see Q.3, Q. 4

2. Discuss the inert character of silver and gold.

Ans: For answer see Q. 6

3. Why are cations smaller and anions are bigger in size than their respective neutral atoms.

Ans: The size of cations is smaller than their parent atoms due to following reasons.

1. Removal of electrons from mentral atoms usually results in the loss of outermost shell.
2. Removal of electrons causes imbalance in picture and electron ratio which causes increase in nuclear charge.

Size anion: The size of anion is greater than parent atom because when an atom gains electrons in the valence shell it increases the element repulsion thus valence shell expands and size of atom increases.

4. *Discuss why hardness and softness of a metal depends upon its metallic bonding.*

Ans: The softness and hardness of metals is directly proportional to the strength of the metallic bond. Stronger the metallic bond, harder will the metal and vice versa.

5. *Give the reaction of sodium with; H_2O, O_2, Cl_2 and H_2.*

Ans: For answer see Q. 4

6. *What are physical properties of calcium metal? Give its uses.*

Ans: 1. Calcium is silvery white metal.

2. It is soft malleable and ductile.
3. It is good conductor of heat and electricity.
4. its melting point is 151°C and boiling point is 1439°C.

Uses: (1) It is used to remove sulphur from petroleum products.

(2) It is used as reducing agent to produce Cr U and Zr.

7. *Write down the chemical properties of the non-metals?*

Ans: For answer see Q. 8

8. *Compare the physical properties of metals and non-metals.*

Ans: For answer see Q. 1(c) and Q. 8

9. *How can you compare the softness and hardness of metals?*

Ans: For answer see Q.3

10. *Give the chemical properties of magnesium and its uses.*

Ans: For answer see Q. 4,5

11. *Write a comprehensive note on the electropositive character of metals?*

Ans: For answer see Q. 2(a)

12. *Compare the ionization energies of alkali and alkaline earth metals.*

Ans: All the elements of alkaline earth metals have high ionization energies as compared to alkali metals as shown in table.

Metal	Atomic Number	IE	Metal	Atomic Number	IE_1	IE_2
Li	3	520	Be	4	899	1787
Na	11	496	Mg	12	738	1450
K	19	419	Ca	20	590	1145
Rb	37	403	Sr	38	549	1064
Cs	55	377	Ba	56	503	965

Low ionization energies of alkali metals make them more reactive than alkaline earth metals.